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ABSTRACT

ChemSource is designed as a strategy to help preservice and inservice high school chemistry teachers promote student learning more effectively. Its major premise is that well-designed laboratory investigations are an important avenue for cultivating student interest, engagement, and meaningful learning in chemistry. The SourceBook component of ChemSource is a resource containing 36 content modules and additional supporting materials which aim at providing specific teaching tips and the best available instructional ideas and information gleaned from experienced, successful chemistry teachers. Each content module features: Content-in-a-Nutshell; Laboratory Activities; Demonstrations; Common Misconceptions; Humor; Applications and Implications; Analogies; Reference Resources; History; and Media. General materials on pedagogy, classroom management, safety, teaching chemistry to the disabled, and other useful resources and tabulations precede and follow the content modules. Volume 4 includes the following: Chemistry of Seawater; Separations; Solutions; Stoichiometry; Thermochemistry; Transition Elements; Computer Uses; Chemistry in Your Career; Cross-Referencing Modules; Teaching Chemistry to Persons with Disabilities; Library Resources; Professional Organizations; Small-Scale Chemistry; and Chemical Safety. (JRH)

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SourceBook

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

VOLUME 4

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MODULES CONTAINED IN SOURCEBOOK

VOLUME 1

Preface and Acknowledgements (PREF)
Chemical Pedagogy (PEDA)
User's Guide (USER)
Introductory: General Resource (INTR)
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Alkali Metals (ALKA)
Atomic Structure (ATOM)
Biogeochemical Cycles (BICY)
Chemical Bonding (BOND)
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THE CHEMISTRY OF SEAWATER

A SourceBook Module

Version 2.0 1994

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ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Life on our planet would be impossible without water. Approximately three-quarters of the earth's surface is covered with water, nearly all of which (98%) is non-potable seawater. There are over fifty million billion (5×10^{12}) tons of sodium chloride and numerous other salts dissolved in the oceans. The chemistry of seawater is the chemistry of a huge solution and can be largely understood in terms of solution chemistry. Drawing on previously covered material, this module will demonstrate how concepts such as colligative properties, pH, density, and equilibrium can be applied to the sea.

In many respects (density, pH, and salinity) seawater is similar to human blood, lending credence to the idea held by most scientists that life evolved from the oceans.

PLACE IN THE CURRICULUM

This module, while it stands alone, could be used piecemeal to add to or replace portions of other modules. For instance, the freezing point and melting point activities could be done in the *Solutions* module, and the conductivity measurements described under demonstrations could be used when teaching the *Bonding* module to illustrate ionic substances. In turn, the *Solutions* module contains demonstrations of the Tyndall effect and the relationship between temperature and gas solubility that are suitable for inclusion in a discussion of seawater, and the *Biogeochemical Cycles* module details the carbon dioxide water equilibrium. This module is a natural extension of the module on *Alkali Metals*, since much of the chemistry of seawater revolves around the metal cations Na^+ and K^+.

CENTRAL CONCEPTS

This is not a fundamental or core module so there are no concepts that absolutely "must be covered." Instead, the central theme of this module is that the chemistry of the oceans is the same as that covered in many other units, but applied to the complex system of the sea. The sea, in turn, is one component of the interaction between land, air, and sea. The questions to be asked and partially answered are:

1. What is seawater?
2. What are the physical and chemical properties of seawater?
3. Where do the components of seawater come from?
4. What are the uses of seawater?

RELATED CONCEPTS

1. Acids/bases
2. Equilibrium, particularly LeChatelier's principle
3. Concentration units
4. Electrolytes and the conducting of electricity
5. Solubility
6. Thermochemistry
7. Colligative properties

RELATED SKILLS

1. pH calculations
2. Predictions of equilibrium shifts based on changing concentrations, pH, temperature, *etc.*
3. Unit conversions, particularly applied to concentrations (*e.g.*, molarity, ppm, g/L)

PERFORMANCE OBJECTIVES

After completing this module, students should be able to:

1. Define the components of seawater.
2. Describe and measure some of the important physical and chemical properties of seawater.
3. Discuss some of the important ramifications of the environment and seawater.
4. State some of the uses of seawater.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Note on Laboratory Activities

These two activities represent only a few examples of the types of experiments that can be used to demonstrate an investigative activity in chemistry. The major concepts of scientific measurement and data analysis, chemical reactions, stoichiometry, properties of solutions, acids and bases, oxidation-reduction reactions, and solubilities, to mention a few, are represented in these two activities.

Activity 1: Analysis of Seawater

Introduction

You will use the methods of chemical oceanography to conduct the water analysis in this investigation. Through these activities, you will solidify such concepts as the physical properties of water and pH. The investigation will be in the form of questions, and you are challenged to answer these questions using your knowledge of chemistry and the scientific method.

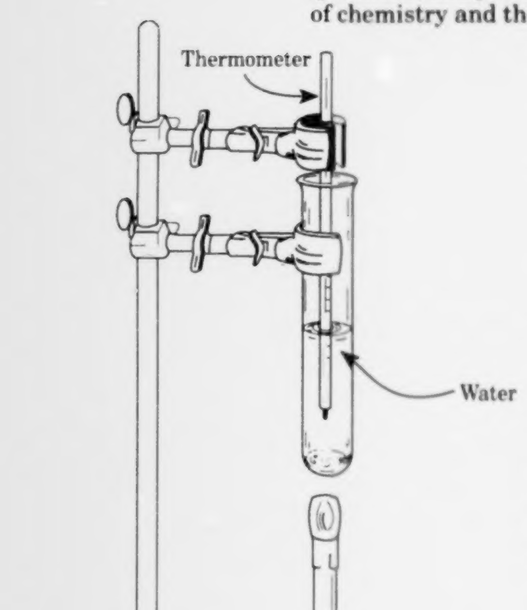


Figure 1. Boiling point apparatus.

Purpose

1. To investigate and compare some of the physical and chemical properties of seawater and distilled water.
2. To make predictions about the physical and chemical properties of seawater and distilled water (based on experiences in the high school chemistry course).
3. To test predictions.
4. To interpret data and draw conclusions based upon experimental results.

Safety

Wear protective goggles throughout the laboratory activity.

Part A. Boiling Point

The boiling point of a liquid is the temperature at which the vapor pressure of the liquid equals the opposing pressure of the atmosphere. The boiling point of water is commonly given as 100 °C at one atmosphere pressure.

Procedure

1. Set up the heating apparatus (Figure 1), measure and record the boiling point of distilled water. (Add boiling chips to prevent bumping.)
2. Record the barometric pressure. Repeat Step 1 with a seawater sample.
3. In a 100-mL beaker, evaporate 20 mL of seawater down to 15 mL. Determine the boiling point of this concentrated seawater using the set-up illustrated in Figure 1.
4. On a day with different weather, repeat Steps 1-3.
5. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

Make a data table containing the boiling points of distilled water, seawater, and concentrated seawater at two different atmospheric pressures.

Implications and Applications

1. Are there any differences between the boiling points of the three substances measured?
2. How did a change in atmospheric pressure affect the boiling points?
3. Are seawater and fresh water affected in the same way by changing atmospheric pressure?
4. Why did increasing the concentration of salt (from Step 3) influence the boiling point?

Part B. Freezing Point

The temperature at which a substance becomes a solid upon cooling is its freezing point, which is identical to its melting point. For water, this temperature is assumed to be 0 °C. In this activity, we will measure the freezing point of water.

Procedure

1. Arrange a thermometer in the test-tube (Figure 2).
2. Cover the bulb of the thermometer with 15 mL water.
3. Pack the test-tube in some chunks of dry ice in a 600-mL beaker.
4. Record the temperature of the water every 30 sec (from the time the tube is just placed in the dry ice), until the temperature stabilizes.
5. Repeat Steps 1-4 using 15 mL seawater.
6. Take 20 mL seawater and evaporate to 15 mL. Measure its freezing point by repeating Steps 1-4.
7. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Make a data table of temperature *vs.* time for all three liquids.
2. Plot the data on a piece of graph paper with the temperature (°C) on the *y*-axis and the time (seconds) on the *x*-axis. Use different symbols for your data points for the three liquids (*e.g.*, x, 0 and D) or different colors to distinguish each curve.

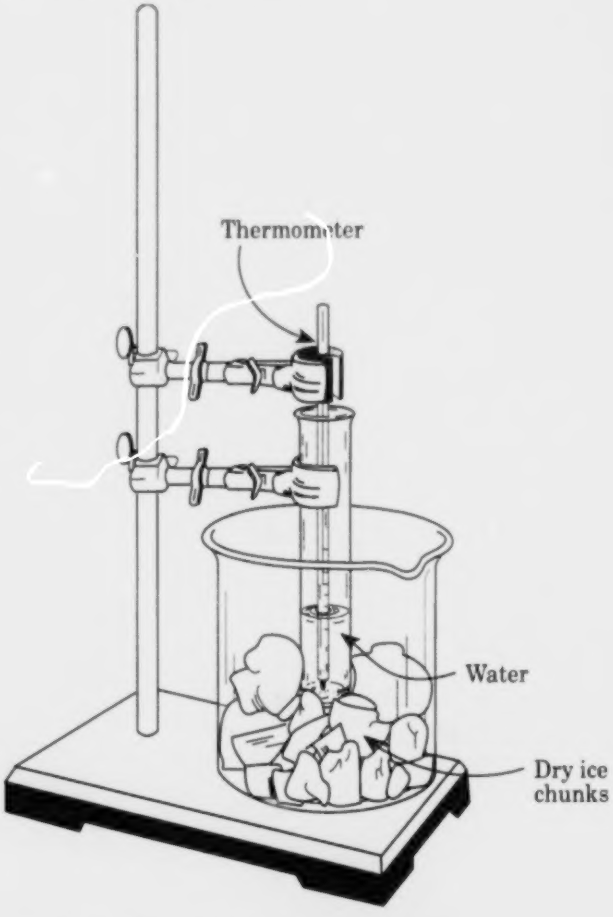


Figure 2. Freezing point apparatus.

Implications and Applications

1. Are there any differences in the freezing points of the three substances measured?
2. Is there any particularly unusual feature of this plot? What does it mean? What caused it?
3. How did increasing the concentration of salt (from Step 6) affect the freezing point of seawater?

Part C. pH of Seawater

The pH of a solution is a measure of its hydronium ion concentration [H_3O^+]. A reading of one on the pH scale represents a high H_3O^+ concentration and a very acidic solution. A very basic or alkaline solution, with low H_3O^+ concentration may read near 14 on the pH scale. A neutral solution will have a pH of 7 on the scale.

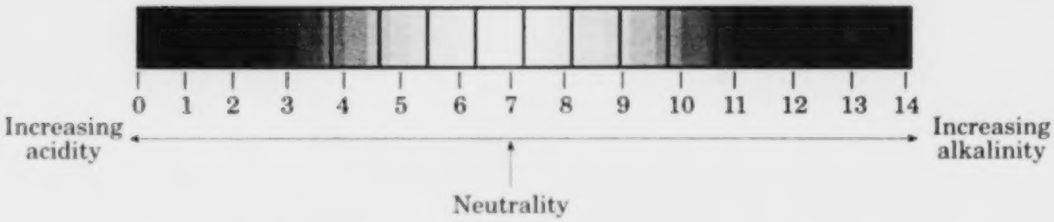


Figure 3. pH ranges.

pH 4.0	Red
pH 5.0	Orange-red
pH 5.5	Orange
pH 6.0	Yellow-orange
pH 6.5	Yellow
pH 7.0	Yellow-green
pH 7.5	Green
pH 8.0	Dark green
pH 8.5	Blue-green
pH 9.0	Blue
pH 9.5	Violet
pH 10.0	Reddish violet

Figure 4. Color chart for Fisher Universal Indicator.

Procedure

Testing distilled water and seawater with acid:

1. Place 50 mL distilled water into a 100-mL beaker. Add 25 drops universal indicator, stir, and record the pH by observing the color of the solution and comparing it with Figure 4.
2. Add one drop 0.1 M HCl, stir, and record the pH.
3. Repeat Step 2 until 20 drops have been added, noting the pH (color) after each drop.
4. Repeat Steps 1-3 using seawater.

Testing distilled water and seawater with base:

5. Repeat Steps 1-3, with distilled water and with seawater, but use 0.1 M NaOH instead of 0.1 M HCl
6. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Collect your data from Steps 1-5 using the following tables.

Drops 0.1 M HCl	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
pH Distilled water																					
pH Seawater																					

Drops 0.1 M NaOH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
pH Distilled water																					
pH Seawater																					

Discussion

Most living organisms can tolerate only slight pH fluctuations near the neutral region of the pH scale. Under open ocean conditions, an effective pH "buffering" system limits seawater pH values to a narrow range between 7.5 and 8.4. This buffering system results from the interaction of dissolved carbon dioxide and water. Much of the carbon dioxide dissolved in seawater combines with water to produce a weak acid, carbonic acid. Normally, carbonic acid dissociates to produce hydrogen ions and bicarbonate or carbonate ions. These reactions can be summarized as shown below. For simplicity, we have represented H_3O^+ as H^+.

$$\underset{\text{Carbon dioxide}}{CO_2(g)} + \underset{\text{Water}}{H_2O(l)} \rightleftharpoons \underset{\text{Carbonic acid}}{H_2CO_3(aq)} \quad (1)$$

$$\underset{\text{Carbonic acid}}{H_2CO_3(aq)} \rightleftharpoons \underset{\text{Hydrogen ion}}{H^+(aq)} + \underset{\text{Bicarbonate ion}}{HCO_3^-(aq)} \quad (2)$$

$$\underset{\text{Bicarbonate ion}}{HCO_3^-(aq)} \rightleftharpoons \underset{\text{Hydrogen ion}}{H^+(aq)} + \underset{\text{Carbonate ion}}{CO_3^{2-}(aq)} \quad (3)$$

If excess hydrogen ions are present, the reactions above proceed to the left. $[H^+]$ decreases, preventing the solution from becoming too acidic. If too few hydrogen ions are present, the reactions above proceed to the right, making more H^+ available and H_2CO_3 is converted to HCO_3^- and CO_3^{2-}.

Normally, the carbonate-bicarbonate buffering system limits large-scale fluctuations in seawater pH. Even so, in restricted water bodies such as tide pools and other areas of limited circulation, the pH of the water can vary enough to strongly affect the organisms living there.

Implications and Applications

1. Compare the buffering action of the distilled water *vs.* that of seawater.
2. Which liquid was the more effective buffer?
3. Can you think of other buffering systems in nature?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Analysis of Seawater

Major Chemical Concept

Part A

The boiling points of distilled water and salt solutions depend on the concentration of solutes and the atmospheric pressure. The greater the ionic concentration, the greater the boiling point (at constant pressure), a result of the vapor pressure lowering caused by the presence of ions in the solution. The change in boiling point can be represented by the equation

$$\Delta T_b = K_b m_{solute}$$

where, ΔT_b is the difference between the boiling point of the solution and that of the pure solvent, K_b is a constant that is characteristic of the solvent and is called the **molal boiling point elevation constant**, and m_{solute} is the *molality* of the solute in the solution.

Part B

The freezing points of distilled water and salt solutions depend on the concentration of solute in the solution. The greater the ionic concentration, the lower the freezing point (at constant pressure), a result of vapor pressure lowering caused by the presence of ions in the solution. The change in freezing point can be represented by the equation

$$\Delta T_f = K_f m_{solute}$$

where, ΔT_f is the difference between freezing point of the solution and that of the pure solvent, K_f is a constant that is characteristic of the solvent and is called the **molal freezing point depression constant**, and m_{solute} is the *molality* of the solute in the solution.

Part C

Seawater has a buffering effect. It resists changes in pH, in contrast to distilled water.

Level

General high school chemistry.

Expected Student Background

Part A

The student should know the use and care of a thermometer and burner.

Part B

The student should know the use and care of a thermometer.

Part C

The student should know how to safely handle acid and base solutions.

Time

Part A

Students should be able to complete *Part A* in one 55-min period if they immediately begin the evaporation of the seawater for Step 3. Step 4 requires an additional 55-min period.

Part B

Students should be able to complete *Part B* in one 55-min period if they immediately begin the evaporation of the seawater for Step 5.

Part C

Students should be able to complete *Part C* in one 55-min period.

NOTE: The teacher can assign one third of the class to do Part A, *one third to do* Part B *and one third to do* Part C. *Students can then share the data on an overhead projector, allowing completion of the laboratory in one day, if desired.*

The teacher can prepare data for Part A, Step 4 *and the concentrated seawater ahead of time to shorten the laboratory time required.*

Safety

Read the *Safety Considerations* in the *Student Version*.

Materials (For 24 students working in pairs)

Part A

Nonconsumables

- 12 Burners and stands
- 12 Thermometers, –10 °C to 110 °C
- 12 Beakers, 100-mL
- 12 Test-tubes, 20- x 150-mm
- Barometer (or call the local weather bureau)

Consumables

- Boiling chips
- Distilled water
- Artificially prepared seawater (chlorinity–19.05 parts per thousand, ppt):
  - *Add to 1 L of distilled water:*
  - 27.53 g Sodium chloride (NaCl)
  - 5.79 g Magnesium chloride hexahydrate ($MgCl_2 \cdot 6H_2O$)
  - 14.17 g Magnesium sulfate heptahydrate ($MgSO_4 \cdot 7H_2O$)
  - 1.82 g Calcium chloride dihydrate ($CaCl_2 \cdot 2H_2O$)
  - 1.23 g Potassium chloride (KCl)
  - 0.27 g Sodium hydrogen carbonate ($NaHCO_3$)

Part B

Nonconsumables

- 12 Ring stands
- 12 Thermometers, –10 °C to 110 °C
- 12 Beakers, 100-mL and 400-mL
- 12 Test-tubes, 20- x 150-mm

Consumables

- Graph paper
- Distilled water
- Seawater (see *Part A* for directions)
- Dry ice (solid CO_2), 6 lb

Part C

Nonconsumables

12 Beakers, 100-mL
12 Dropping bottles, Fisher universal indicator, pH range 4-10
12 Magnetic stirrers (optional)
24 Medicine droppers, or equivalent

Consumables

0.1 M Hydrochloric acid, HCl (4.2 mL 12 M HCl to 500 mL solution)
0.1 M Sodium hydroxide, NaOH (2.0 g NaOH to 500 mL solution)

Advance Preparation

Parts A and B

Prepare:

Artificial seawater or obtain seawater
Concentrated seawater (optional)

Part C

In addition to *Part A and B*, prepare:

0.1 M HCl
0.1 M NaOH

Pre-Laboratory Discussion

1. Students should be instructed to start the evaporation of the salt solution to make the concentrated salt solution first (if it is not being provided by the teacher).
2. Review the concepts of pH, buffers and colligative properties.
3. Discuss supercooling and how it might appear on a cooling curve.
4. Review proper graphing technique including title, axes labels, proper sizing of units on each axis.
5. Warn the students not to use the thermometer as a stirring rod because the fragile bulb at the end of the thermometer could break.
6. For *Part C*, if you have buffer solutions from pH 4 through pH 10, it would be helpful if students could see the actual universal indicator colors shown in Figure 4.

Teacher-Student Interaction

Circulate in the laboratory. Ask questions regarding expected results.

Post-Laboratory Discussion

1. Have the students place their freezing point and boiling point results on the board or overhead projector and determine the average class values.
2. Discuss why there is a range of values (difference in equipment, rate of heating, and rate of cooling).
3. Discuss the shapes of the cooling curves from *Part B*.
4. Discuss other natural buffering systems, like blood.

Anticipated Student Results

Part A

	Boiling point, °C
Distilled water	100.0 ± 0.5
Seawater	102.5 ± 0.5
Conc. seawater	103.0 ± 0.5

Figure 5. Student data table. Boiling points at pressure of 757 mm Hg.

Part B

Temperature °C	Time seconds
24	0
17	30
14	60
10	90
9	120
8	180
5	210
4	240
3	270
2.5	300
2	330
–3	360
1	390
0	420
0	450
0	480

Figure 6. Temperature vs. time for distilled water.

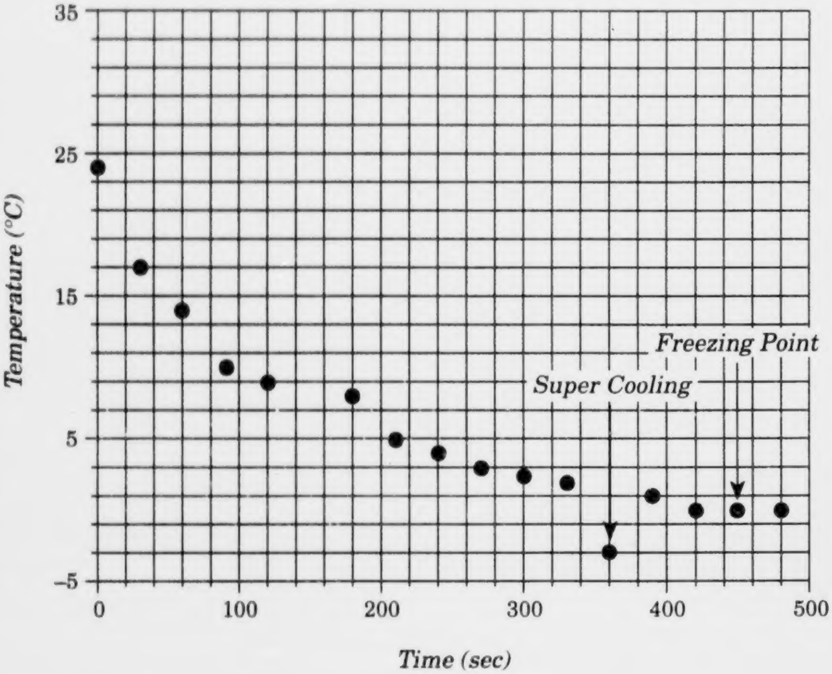


Figure 7. Freezing point curve for distilled water.

Temperature °C	Time seconds
23	0
16	30
12	60
5.5	90
2	120
1	150
–0.5	180
–1.5	210
–2	240
–2.5	270
–3	300
–3.5	330
–4	360
–5	420
–6	450
–6.5	480
–7	510
–8	540
–8	570
–8	600

Figure 8. Temperature vs. time for seawater.

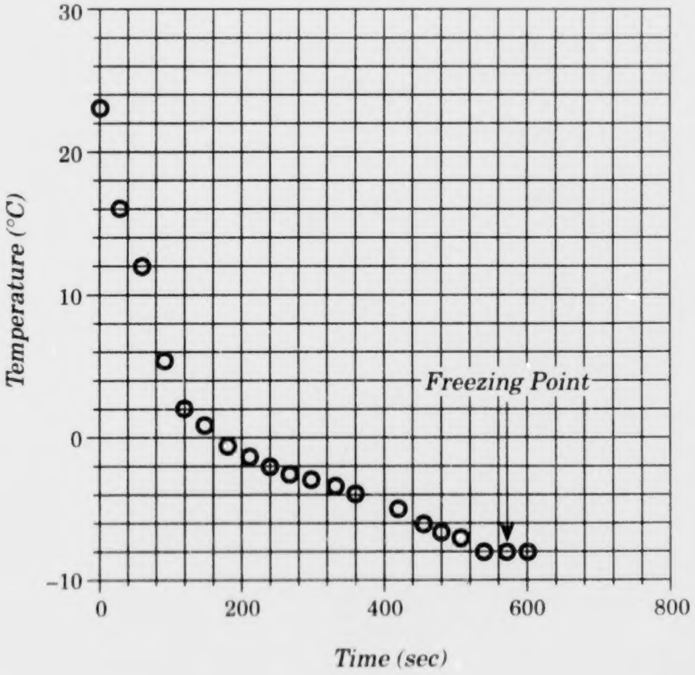


Figure 9. Freezing point curve for seawater.

Temperature °C	Time seconds
28.5	0
16	30
10	60
8	90
5.5	120
3.5	150
0	180
–5	210
–7	240
–8.5	270
–9	300
–9	330
–9	360

Figure 10. Temperature vs. time for concentrated seawater.

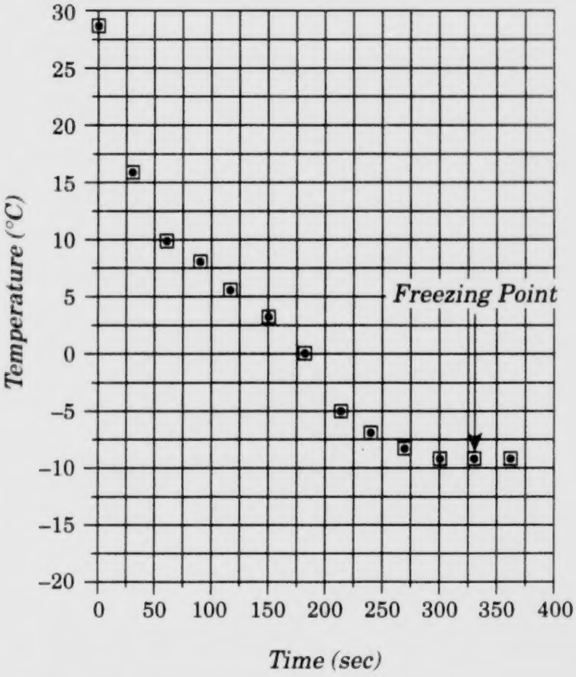


Figure 11. Freezing point curve for concentrated seawater.

Salinity (ppt)	Freezing point (°C)
0	0
10	–0.53
20	–1.08
30	–1.63
35	–1.91

Figure 12. Freezing point of seawater as a function of salinity.

Using the data from Figure 12 and their experimental freezing points, students can infer the concentrations of dissolved salts in their distilled water, seawater and concentrated seawater in parts per thousand.

Part C

Drops of 0.1 M HCl	pH Distilled water	pH Seawater
0	7.5	7.5
1	6.5	7.0
2	5.0	7.0
3	4.0	7.0
4	<4.0	7.0
5	<4.0	7.0
6	<4.0	7.0
7	<4.0	7.0
8	<4.0	6.5
9	<4.0	6.5
10	<4.0	6.5
11	<4.0	6.5
12	<4.0	6.5
13	<4.0	6.5
14	<4.0	6.5
15	<4.0	6.5

Figure 13. pH values of distilled water and seawater on addition of 0.1 M HCl.

Drops of 0.1 M NaOH	pH Distilled water	pH Seawater
0	7.5	7.5
1	9.0	8.0
2	9.5	8.5
3	10.0	8.5
4	>10.0	8.5
5	>10.0	8.5
6	>10.0	8.5
7	>10.0	8.5
8	>10.0	8.5
9	>10.0	8.5
10	>10.0	9.0
11	>10.0	9.0
12	>10.0	9.0
13	>10.0	9.0
14	>10.0	9.0
15	>10.0	9.0

Figure 14. pH values of distilled and seawater on addition of 0.1 M NaOH.

Answers to Implications and Applications

Part A

1. The boiling point of seawater increases as the salt concentration increases:

   Conc. seawater > Seawater > Distilled water
2. The boiling points increase slightly as the pressure increases.
3. Yes.
4. Increasing the salt concentration increases the boiling point, by molal boiling point elevation, a colligative property (see *Solutions* module).

Part B

1. Yes. Distilled water > Seawater > Conc. seawater.
2. The temperature stabilizes as soon as ice starts to form in the test-tube.
3. The freezing point decreases as the salt concentration increases, by molal freezing point depression, a colligative property (see *Solutions* module).

Part C

1. The pH changes rapidly with the addition of HCl or NaOH to distilled water. Addition of even large quantities of HCl or NaOH to seawater changes the pH only slightly.
2. Seawater.
3. One example is the buffering of blood by a CO_2/HCO_3^-. (There are also other buffers in the blood.)

Post-Laboratory Discussion

A discussion of supercooling will probably be in order after the students have plotted their results for *Part B*.

The teacher may wish to discuss the importance of the buffering effect of the HCO_3^-/CO_3^{2-} in the oceans and the environment.

Assessing Laboratory Learning

1. What is a buffer solution? *[A solution that resists changes in its pH when either an acid or a base is added.]*
2. Give an example of buffer solutions that occur in nature. *[Blood and seawater (any solution that resists changes in pH).]*
3. As the concentration of a salt in a solution is increased, what happens to the freezing point. Explain. *[The freezing point decreases. The depression is due to the colligative properties of the system (the solution has a lower vapor pressure than that of pure water).]*
4. As the concentration of a salt in a solution is increased what happens to the boiling point. Explain. *[The boiling point increases. The elevation is due to the colligative properties of the system (a nonvolatile solute elevates the boiling point of the solution because the vapor pressure of the liquid is decreased).]*
5. If a nonvolatile soluble substance were added to a liquid other than water, predict how its boiling point and freezing point be affected. Why? *[The boiling point would be elevated and the melting point would be depressed, but not to the same degree as water since the* K_f *and* K_b *values are different for each solvent.]*

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Determination of the Salinity of Seawater

Introduction

This method involves chemical volumetric analysis. A silver nitrate ($AgNO_3$) solution of known concentration is used to precipitate out the chlorides in a seawater sample. The procedure is complicated by the fact that the dense white precipitate, silver chloride (AgCl), does not settle rapidly. It is impossible to tell when all the chlorides have been removed from the seawater unless an indicator is used to signal complete precipitation by a visual color change. The indicator in the analysis is chromate ion. When all of the chloride ion is exhausted, the chromate ion reacts with silver ions and produces silver chromate, which is red. The instant a permanent orange tinge appears in the solution (one that doesn't vanish with mixing), the addition of silver nitrate is stopped. The final solution color should look like that of orange juice.

The chemical reactions are:

$$Cl^-(aq) + Ag^+(aq) \rightarrow \underset{\text{(white)}}{AgCl(s)} \qquad (1)$$

$$2Ag^+(aq) + CrO_4^{2-}(aq) \rightarrow \underset{\text{(red-brown)}}{Ag_2CrO_4(s)} \qquad (2)$$

The titration is carried out in a 150-mL Erlenmeyer flask, using a white background and good lighting, but away from direct sunlight. Constant but gentle stirring is essential, since the silver chloride tends to form clots, which trap some of the reagents. A magnetic stirrer is very helpful. The silver nitrate is added until the faint reddish or orange tinge remains after stirring for 45 s. At this point one drop should give the desired end point. Observable flocs of silver chloride will obscure the end point, since the solutions that they contain in their interstices will leak out slowly and cause the end point to be impermanent. A stirring rod left permanently in the flask, especially if a magnetic stirrer is not used, will help to break up these curds.

Purpose

1. To determine the chlorinity (and thus the salinity, since they are directly related) of a sample of seawater.
2. To use the titration technique with a reaction different from an acid/base reaction.

Safety

Wear protective goggles throughout the laboratory activity. Avoid getting $AgNO_3$ on hands since it stains the skin.

Procedure

1. With a pipet deliver 5.0 mL seawater into a 150-mL Erlenmeyer flask.
2. Using a 10-mL graduated cylinder, add 5.0 mL potassium chromate indicator solution, and begin to add the silver nitrate solution from a buret while constantly stirring.
3. When the solution turns orange, stop adding silver nitrate. The color should persist for less than 20 sec. Continue adding silver nitrate dropwise, mixing all the while, until orange color persists. Record the volume of silver nitrate added.
4. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

From a simple algebraic proportion one can determine the chlorinity of the seawater because it has been shown experimentally that the salinity is a constant 1.80655 times higher than the chloride level. Thus, the salinity (S) can be calculated (using the equation S ppt = 1.80655 Cl ppt; the chlorinity (Cl) of the artificial seawater should be about 19.05 ppt).

1. What is the number of moles of silver used in the titration ($AgNO_3$ is 0.2184 M)?
2. Determine the number of moles of chloride in the sample.
3. Determine the grams of chloride present in the sample.
4. What is the chlorinity of 10 g of sample? (The density of seawater is about 1.05 g mL^{-1}).
5. What is the salinity of the sample in ppt?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Determination of the Salinity of Seawater

Major Chemical Concept

Quantitative determination of the chloride concentration in a sample of seawater.

Level

General high school chemistry

Expected Student Background

Students should have done titrations and titration calculations.

Time

55 min

Safety

Read the *Safety Considerations* in the *Student Version*.

Materials (For 24 students working in pairs)

Nonconsumables

12 Burets
12 Buret stands and clamps
12 Graduated cylinders, 10-mL
12 Erlenmeyer flasks, 150-mL
12 Magnetic stirrers (optional)
12 Pipets, transfer, 10-mL
2 Storage bottles, 1-L
1 Volumetric flask, 1-L

Consumables

Distilled water (in wash bottle)
Indicator solution (3.5 g, potassium chromate, K_2CrO_4, reagent grade, per liter solution)
0.2184 M Silver nitrate, $AgNO_3$ (37.11 g $AgNO_3$, reagent grade, to 1.000 L in a 1-L volumetric flask. Store solution in a tightly capped brown glass bottle.)

Advance Preparation

Prepare:

Potassium chromate (observe safety precautions)
Silver nitrate (observe safety precautions)

Pre-Laboratory Discussion

1. Demonstrate the proper titration techniques and remind students that silver nitrate will stain their hands.
2. Discuss the reaction that gives the red-brown color

   a. $$Cl^-(aq) + Ag^+(aq) \rightarrow \underset{\text{white}}{AgCl(s)}$$

   b. $$2Ag^+(aq) + CrO_4^{2-}(aq) \rightarrow \underset{\text{red-brown}}{Ag_2CrO_4(s)}$$

   and demonstrate a titration showing the students the final end point color.
3. Review the proper use of a transfer pipet.

Teacher-Student Interaction

Circulate in the laboratory. Ask questions comparing this titration with previously done acid-base titrations. Ask why it is not important to use an exact amount of indicator solution.

Post-Laboratory Discussion

Part A

Discuss differential solubility. Since the $Ag_2CrO_4(s)$ is much more soluble than $AgCl(s)$, it is only after almost all of the $AgCl(s)$ is precipitated that the red-brown end point be reached with the precipitation of $Ag_2CrO_4(s)$.

Anticipated Student Results

Trial	Volume of $AgNO_3$	Concentration of $AgNO_3$
1	28.05 ml	0.2184 M

Answers to Data Analysis and Concept Development

1. $(0.2184\ M\ AgNO_3)(0.02805\ L) = 0.006126\ mol\ AgNO_3 = 0.006126\ mol\ Ag^+$
2. $mol\ AgNO_3 = mol\ Cl^- = 0.006126\ mol$
3. $(mol\ Cl^-)\left(\frac{35.45\ g\ Cl^-}{1\ mol\ Cl^-}\right) = 0.2172\ g\ of\ Cl^-$
4. $\left(\frac{mass\ Cl^-}{mass\ 10\ mL\ seawater}\right)(1000) = Chlorinity\ (ppt)$

   $\left(\frac{0.217\ g\ Cl^-}{10.5\ g\ seawater}\right)(1000) = 20.7\ ppt$
5. $(1.80655)\ chlorinity(ppt) = Salinity\ (ppt)$

   $(1.80655)(20.7\ ppt) = 37.4\ ppt$

Post-Laboratory Discussion

Discuss solubility. Since the $Ag_2CrO_4(s)$ is much more soluble than $AgCl(s)$, it is only after almost all of the $AgCl(s)$ is precipitated that the red-brown end point is reached with the precipitation of $Ag_2CrO_4(s)$.

Assessing Laboratory Learning

1. If 24.0 mL of 0.250 M $AgNO_3$ is required to reach the end point of the seawater solution ($CrO_4^{2-}(aq)$ reaction with $Ag^+(aq)$ to give a red-brown precipitate), determine the number of moles of Cl^- ions present.

   [mol of $AgNO_3$ = (volume of $AgNO_3$) (molarity of $AgNO_3$)
   mol of $AgNO_3$ = (0.0240 L) (0.250 M) = 0.00600 mol of $AgNO_3$
   mol of $AgNO_3$ = mol of Cl^- = 0.00600 mol Cl^-]

2. Determine the grams of chloride present in the sample.

   $[(0.00600\ mol\ of\ Cl^-)\frac{35.45\ g\ Cl^-}{1\ mol\ Cl^-} = 0.213\ g\ of\ Cl^-]$

3. What is the chlorinity of a 10.0 mL sample? (The density of the seawater is determined to be 1.05 g/mL.)

$$\left[\frac{\text{Mass of } Cl^-}{\text{Mass of 10.0 mL of seawater}}(1000) = \text{Chlorinity (ppt)}\right.$$

$$\left.\frac{0.213 \text{ g } Cl^-}{(1.05 \text{ g/mL})(10.0 \text{ mL})}(1000) = 20.3 \text{ ppt}\right]$$

4. What is the salinity of the sample? (1.80655) Chlorinity (ppt) = Salinity (ppt)

   [(1.80655) Chlorinity (ppt) = Salinity (ppt)
   (1.80655) (20.3 ppt) = 36.7 ppt]

CAUTION: Use appropriate safety guidelines in performing demonstrations.

DEMONSTRATIONS

Demonstration 1: Comparison of the Electrical Conductivity of Distilled Water with Tap Water and Seawater

Purpose

The electrical conductivity of distilled (or deionized) water is compared with tap water and seawater as one indication of the amount of salt present. (The *Bonding* module describes the pertinent apparatus.)

Materials

9-volt Battery conductivity apparatus
Distilled water
Seawater

Safety

Use normal safety precautions.

Procedure

Test electrical conductivity of distilled water and seawater with the apparatus.

Extensions

Discuss why it is inadvisable to go swimming during a lightning storm.

Demonstration 2: Buoyancy and Density

Purpose

Part A simply demonstrates what buoyancy is and how we recognize it. *Part B* shows the relationship between density and buoyancy.

Materials

Part A

Beaker, 2-L or larger (or clear plastic pail)
Ice
Water, 2 L
Canned soft drinks, sweetened and sugar-free

Part B

2 Graduated cylinders, 1-L
Water (tap water or deionized water), 2 L
400 g Sodium chloride, NaCl, saturated solution, per L solution
2 Eggs (raw)

Safety

Use normal safety precautions.

Procedure

Part A

Fill a very large (2 L or larger) beaker or clear plastic pail with water and a little ice. Drop in one can of Pepsi and one can of Diet Pepsi (or Coke and Diet Coke). The sugared solution will sink to the bottom while the artificially sweetened solution, which is much less dense, will float at or near the top of the container.

Part B

Fill one graduated cylinder with tap water or deionized water; in the other place 250 mL saturated NaCl solution. Now, carefully pour more water down the side of the cylinder without stirring until the cylinder is nearly full. Now drop an egg into each container. In pure water, the egg falls. In the salt solution, the egg will remain suspended somewhere in the bottom half of the cylinder where the density of the solution equals that of the egg.

Remarks

One way to introduce this demonstration is to ask if any students have ever swum in the Great Salt Lake or the Dead Sea. If so, they can attest to the difference in buoyancy between those very salty "oceans" and fresh water.

Maybe students will have noticed the phenomenon illustrated in *Part A* at picnics where large trash cans or barrels are often filled with ice and canned drinks. If there is freedom of movement, the sugared drinks will be on the bottom and diet drinks on top. The greater the sugar content, the deeper the can sinks.

Demonstration 3: Osmotic Pressure

Purpose

In this activity, students will observe increased osmotic pressure as a saturated solution of sodium chloride is diluted by distilled water passing through a membrane causing the water level of a thistle tube to rise.

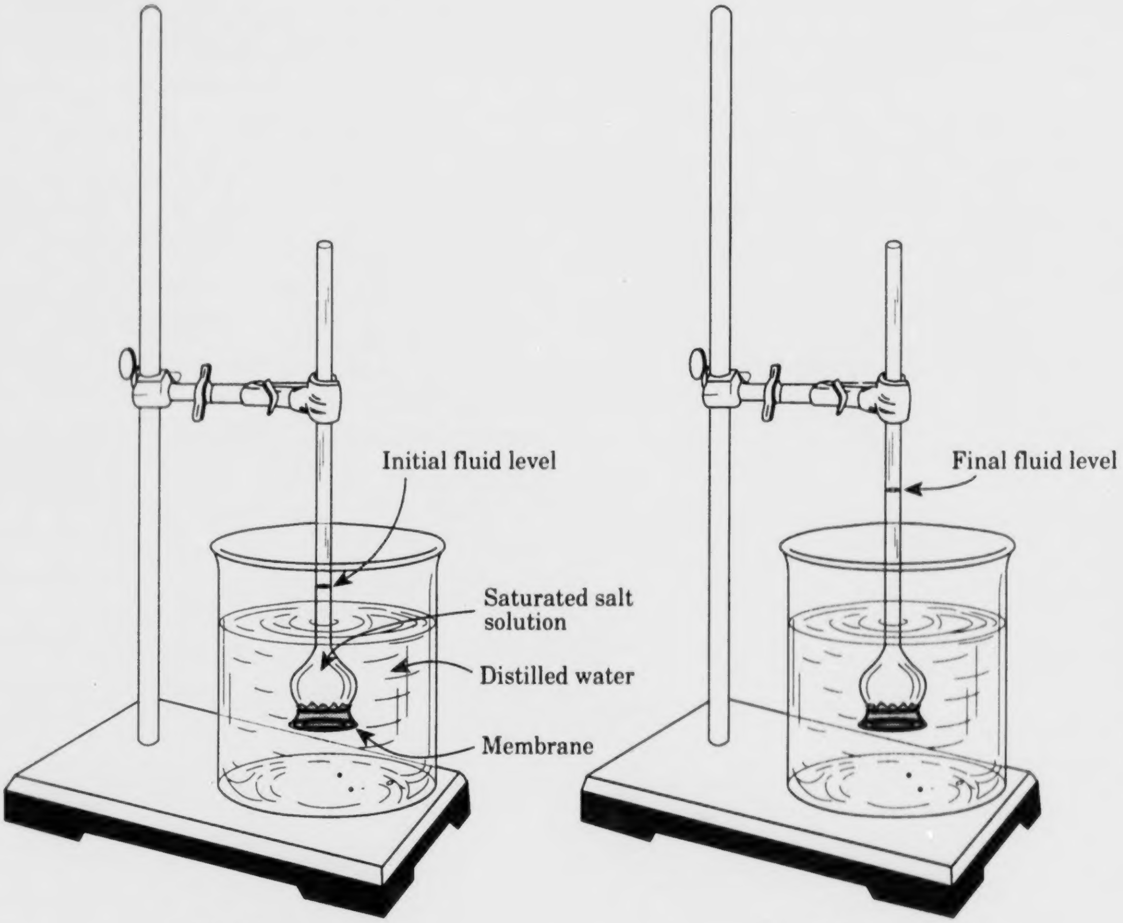
Materials

Ring stand and clamp
Beaker, 400-mL and 150-mL
Semipermeable membrane (dialysis tubing or cellophane)
Saturated salt water (40 NaCl per 100 mL solution)
Thistle tube or similar tubing

Procedure

Fill 400-mL beaker approximately 3/4 full of distilled water. Put saturated NaCl solution inside of thistle tube (which has membrane over the large open end) until saturated NaCl solution is just above the distilled water level in the beaker (see Figure 15). Mark initial level on Day 1 and mark final level the next day. *[NOTE: To fill the thistle tube, place piece of cellophane in a 150-mL beaker and depress it to the bottom with thistle tube. Remove thistle tube and fill depression with NaCl (sat.). Replace thistle tube and allow to fill with solution. Secure cellophane around bottom of thistle tube with thread or rubber band.]*

The distilled water passed through the membrane and diluted the saturated sodium chloride solution. As this happened, pressure increased, the fluid level inside of the thistle tube increased, and the fluid level in the thistle tube rose.


Initial fluid level
Saturated salt solution
Distilled water
Membrane
Final fluid level

First Day

Second Day

Figure 15. Osmotic pressure apparatus.

Extensions

1. Evaporate some seawater and look at the remaining solids under a microscope.
2. Again, using a microscope, look at plant cells and what happens when they are immersed in distilled water *vs.* salt water.

Remarks

Osmosis is used to desalinate water. In light of the severe water shortage in certain parts of the world, there is great interest in using osmosis to purify water.

GROUP AND DISCUSSION ACTIVITIES

Counterintuitive Examples

1. Water can be purified by passing through a filter. *[Some impurities or substances found in water can be filtered. Suspended solids can be removed with filter paper; some organic compounds can be absorbed on pulverized charcoal. Soluble salts are not so easily removed.]*

2. All salts dissolve more in warm water than in cold water. *[While this fact is true for many salts, it is not universal. For example, when NH_4NO_3 dissolves in water, the reaction is endothermic, meaning that heat is absorbed during the reaction. The surrounding air and solution get colder as they supply this heat.*

   $$Heat + NH_4NO_3(s) + H_2O(l) \rightarrow NH_4^+(aq) + NO_3^-(aq)$$

   According to LeChatelier's Principle raising the temperature of the system, which is adding more of the reactant heat, will cause more NH_4NO_3 to dissolve.

   In contrast, when $MgSO_4$ dissolves in water, the opposite effect is noted; the surroundings get warmer because heat is produced.

   $$MgSO_4(s) + H_2O(l) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq) + Heat$$

   In this case, raising the temperature shifts the equilibrium to the left so that less $MgSO_4$ dissolves in hot water than in cold.

   As a side note, mention to students that NH_4NO_3 is the salt used in many first-aid cold packs, while $MgSO_4$ is used in first-aid hot packs.]

3. Gases dissolve better in warm water than cold. *[Open a cold can of soda and one at room temperature. Which one loses its carbonation more quickly? (The warmer one because CO_2 is less soluble at higher temperatures.)]*

Pictures in the Mind

1. Distribution of the world's water supply (see *Appendix*).
2. Seawater contains an average of 35,000 parts per million of dissolved solids. In a cubic mile of seawater, weighing 4.7 billion tons, there are about 165 million tons of dissolved matter, mostly chlorine and sodium (see *Appendix*).

TIPS FOR THE TEACHER

Language of Chemistry

boiling point temperature at which the vapor pressure of a liquid is equal to the external pressure.

boiling point elevation a colligative property in which the boiling point of a solution is raised above that of the pure solvent according to the number of solute particles present.

buffer a solution that resists change in pH when either OH^- or H^+ is added.

chlorinity the total amount of dissolved halide ions (Cl^-, F^-, Br^-, I^-) present in a solution such as seawater; usually expressed in mol L^{-1}, ppt or ppm.

colligative properties physical properties of a solution (vapor pressure lowering, boiling point elevation, freezing point depression, osmotic pressure) that are different for a solution than for a pure solvent, according to the number of solute particles, and not their identities.

freezing point the temperature where the solid and liquid phase of a substance coexist; also called melting point.

freezing point depression a colligative property in which the freezing point of a solution is lowered below that of the pure solvent according to the number of solute particles present.

melting point see freezing point.

osmosis the flow of solvent into a solution through a semi-permeable membrane.

osmotic pressure a colligative property—the pressure which must be applied to a solution to stop osmosis.

parts per billion (ppb) a unit of concentration. One ppb is one part of solute per billion parts of solution. For example, if an air sample contained 1 ppb Hg vapor then there is 1 mL Hg(g) in 10^9 mL air or 1 mL Hg in 10^6 L of air.

parts per million (ppm) one ppm is one part of solute per million parts of solution.

parts per thousand (ppt) one ppt is one part of solute per thousand parts of solution.

salinity total concentration of dissolved material (all ions), usually expressed in mol L^{-1} or ppt.

Common Student Misconceptions

Gold can be mined from seawater. While seawater does contain a small amount of gold (0.011 ppb), it is not economically feasible to separate it. Have students calculate how much gold there is in the oceans, which contain approximately 3×10^{19} L of water (see *Counterintuitive Examples*).

Decision Making

1. How well must sewage be treated before dumping in the ocean? Need it be pure water when it leaves the wastewater treatment plant or is some amount of dissolved material more ecologically sound?
2. Effect of global warming on the ocean. As scientists worry about increasing CO_2 and CH_4 concentrations in the atmosphere, what are the possible effects on the ocean?
3. Additional CO_2 in the atmosphere can lead to greater quantities dissolved in seawater. What effects might occur?

History: On the Human Side

In the Bible, the cities of Sodom and Gomorrah are associated with sin and destruction. Traditionally, these cities were located on the large plain south of the Dead Sea, where grotesque salt pillars (a result of the rapid evaporation of the briny waters) can still be seen. We remember that as Lot fled the destruction of these cities, his wife looked back and turned into a pillar of salt. What do you think about this story and its geographical setting?

Humor: On the Fun Side

1. The children of the Bluestone family were playing a game of hide-and-go-seek, and all were enjoying the game save the youngest child who was always the first to be discovered. When asked why this was so, he replied, "I guess I just don't know how to hydrate."

   CHEM 13 NEWS, December 1984, p. 824
2. CH_2O is the formula for seawater.
3. Message on a T-shirt:

   Front: What's $(H_2O)_4$?
   Back: Drinking.
4. $\frac{[NaCl(aq)]^2}{C_7}$ "Sailing, sailing, over the seven seas."
   (Saline, saline, over the seven C's.)
   CHEM 13 News, November 1976, p. 1107

5. **AN ODE TO WATER** *(To be sung to the Mickey Mouse Club tune)*

   What's the solvent molecule, that's found in you and me?
   DoubleU-A-T-E-R, as bent as it can be.

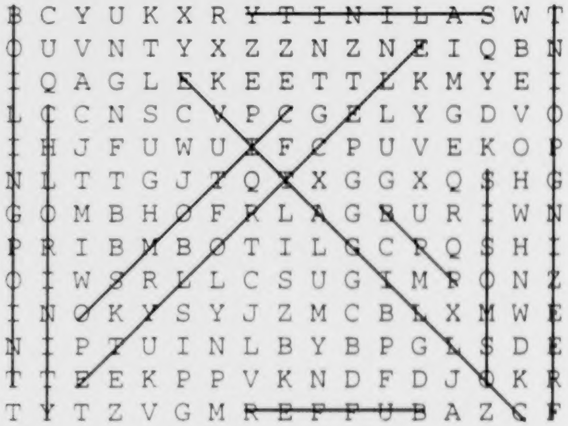
   If it could be linear, then gases we would be.
   DoubleU-A-T-E-R, just right for you and me.

   Hydrogen,
   Oxygen,
   You give it just the right polarity.

   We applaud its structure, with its subtle symmetry.
   DoubleU-A-T-E-R, as bent as it can be.

   CHEM 13 NEWS, May 1988, p. 2

6. Word Search (see *Appendix* for master copy)



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B C Y U K X R Y T I N I L A S W T
O U V N T Y X Z Z N Z N E I Q B N
I Q A G L E K E E T T L K M Y E I
L C C N S C V P C G E L Y G D V O
I H J F U W U I F C P U V E K O P
N L T T G J T Q T X G G X Q S H G
G O M B H O F R L A G B U R I W N
P R I B M B O T I L G C P Q S H I
O I W S R L L C S U G I M P O N Z
I N O K Y S Y J Z M C B L X M W E
N I P T U I N L B Y B P G L S D E
T T E E K P P V K N D F D J O K R
T Y T Z V G M R E F F U B A Z C F
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   Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

   1. Total concentration of dissolved material (all ions), usually expressed in mol/L.
   2. Flow of solvent into a solution through a semi-permeable membrane.
   3. Total amount of dissolved halide ions present in a solution such as seawater.
   4. Solution that resists change in pH when either OH^- or H^+ is added.
   5. Adjective for physical properties of a solution that are different from those of a pure solvent and depending on the number of particles dissolved.
   6. Pressure that must be applied to a solution to stop osmosis.
   7. Temperature where the solid and liquid phase of a substance coexist (2 words).
   8. Temperature at which the vapor pressure of the liquid is equal to the external pressure (2 words).
   9. Unit of concentration abbreviation meaning one part of solute per billion parts of solution.
   10. Substance whose solution in water conducts electricity.

   Answers: 1. SALINITY 2. OSMOSIS 3. CHLORINITY 4. BUFFER 5.COLLIGATIVE 6. OSMOTIC 7. FREEZING POINT 8. BOILING POINT 9. PPB 10. ELECTROLYTE

7. See cartoons at end of module.

MEDIA

1. The *World of Chemistry* videotapes "No. 12: Water," "No. 17: The Precious Envelope," and "No. 18: The Chemistry of the Earth," are particularly appropriate for this module. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. Videodiscs published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    a. "Water's Structure," "Detergent Action," "The Hydrologic Cycle" and "Acid Rain and Limestone," four chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.
    b. "Water and Ice," "A Solution with Water," "Alcohol Dissolves in Water," and "Water and Oil," four chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

Links/Connections

WITHIN CHEMISTRY

1. See module on *Solubility and Precipitation*.
2. The volume of the earth's oceans is 1.4×10^{21} L, with a concentration of sodium chloride of 31.3 g per kg of seawater. Total number of moles of sodium chloride in the ocean is 7.5×10^{20}, which is equivalent to 4.4×10^{22} g or 4.9×10^{16} tons.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Biology (saline solutions and solution concentrations)
Geology
Economics
Physics (physical separations)
Oceanography

Constituent		g/kg Seawater	Percent by weight
Chloride	Cl^-	18.980	55.04
Sodium	Na^+	10.556	30.61
Sulfate	SO_4^{2-}	2.649	7.68
Magnesium	Mg^{2+}	1.272	3.69
Calcium	Ca^{2+}	0.400	1.16
Potassium	K^+	0.380	1.10
Bicarbonate	HCO_3^-	0.140	0.41
Bromide	Br^-	0.065	0.19
Boric acid	H_3BO_3	0.026	0.07
Strontium	Sr^{2+}	0.013	0.04
Fluoride	F^-	0.001	0.00
	Total	**34.482**	**99.99%**

Dissolved constituents of 1 kilogram of seawater (Chlorinity = 19%)

Constituent		Percent by weight
Carbonate	CO_3^{2-}	35.15
Calcium	Ca^{2+}	20.39
Sulfate	SO_4^{2-}	12.14
Silicon Dioxide	SiO_2	11.67
Sodium	Na^+	5.79
Chloride	Cl^-	5.68
Magnesium	Mg^{2+}	3.41
Oxides	Fe_2O_3, Al_2O_3	2.75
Potassium	K^+	2.12
Nitrate	NO_3^-	0.90
	Total	**100.00%**

Dissolved constituents of river water (average)

TO THE CONTEMPORARY WORLD

Personal

1. Kelp (seaweed) is harvested to provide many products. The most important of these is probably algin, or sodium alginate, a polysaccharide with gelling properties. It finds uses in ice cream and soft drinks.
2. The similarities between seawater and blood:
   - Both composed of mostly water.
   - Both have similar densities.
   - Both contain similar percentages of dissolved salts.
   - Both have similar pH values.
   - Both are slightly basic.

Community

Obviously, teachers in communities on or near an ocean (or the Great Salt Lake) can plan many field trips and take advantage of experts in the area. Teachers in inland cities can use museums and museum directors as resources. Marine biologists and chemists are hired by aquariums (Seaworld and Disney Epcot Center, for example) to maintain balanced ecosystems in display tanks.

Societal (Science/Technology/Society; Current Events)

1. Increasingly, the desalinization of seawater looks attractive as a source of fresh water for populations of arid regions.
2. Sea salts are "farmed" in many areas by drawing the seawater into evaporating ponds.
3. The sea is a source for many very important chemical separations. It is the largest source of Br^- salts, which are used to make Br_2.
4. A discussion of some methods of desalination that are currently used:
   - **Distillation**. In a closed system, evaporation can remove salt from seawater. Upon condensation this fresh water can be isolated from the seawater and collected.
   - **Freezing**. Because salt interferes with the hydrogen bonding that occurs as water freezes, water will crystallize first where the hydrogen bonding is not inhibited. Therefore, as ice freezes it excludes the salt from the crystal, effectively purifying itself.
   - **Reverse osmosis**. A semipermeable membrane (a membrane permeable to H_2O molecules but not to ions or larger molecules) can separate fresh water from seawater. The fresh water, by osmosis, passes into the seawater and dilutes it. The level on the seawater side rises since more H_2O molecules enter this side than leave. In order to force more H_2O molecules back into the fresh water side an external force in excess of the osmotic pressure of normal seawater (24.8 atm) is exerted on the seawater side. More fresh water is forced back into the fresh water side than diffused into the seawater. The net result: production of fresh water.
   - **Electrodialysis**. Very similar to reverse osmosis. A semipermeable membrane selectively isolates the salt from the fresh water. An electrical potential difference speeds the process by attracting the ions.

5. Hydrocarbons in the form of oil and natural gas are being produced from wells on the continental shelves in many areas of the world. Some of the issues raised by such drilling:
    - The size of the oil reservoir.
    - Active earthquake faults.
    - Prevailing winds and currents in the area of a drilling platform.
    - Ecological, social, and economic impacts of drilling activities on the community.
    - Impact of drilling with current oceanic resources such as fishing and recreation.
6. Manganese nodules on the ocean floor contain large amounts of iron and manganese.
7. Phosphate compounds have accumulated on submarine terraces on many continental shelves.
8. Potential gold and diamond deposits lie at the mouths of large rivers.
9. Muds near hot springs on the ocean ridges are rich in metallic minerals.
10. Energy can be derived from the sea by tides, waves, and thermal differences.
11. Major and minor ions in seawater are present in living systems in the same relative concentrations. Does this give any credence to the hypothesis that life originated in the oceans?

Extensions and Projects

1. Discuss the chemistry and critical features of aquarium water. Why do some fish live only in fresh water while some live only in salt water? What physiological adaptations do salmon make as they move from fresh water to seawater?
2. The hydrothermal vents in the trenches of the ocean exhibit some fascinating chemistry that could be investigated.

References

Module drafted by Joseph Baron, Marvin Blevins, and Barbara Sawrey, the California team.

American Public Health Association, American Water Works Association, and Water Pollution Control Federation. (1989). *Standard methods for the examination of water and wastewater* (17th Ed.). Washington, DC: Author.

Goldman, M. (1993, September). The Dead Sea as a source of life – The first plague. *Chem 13 News*, p. 35.

Ingmanson, D. E., and Wallace, W. J. (1989). *Oceanography: An introduction* (4th Ed.). Belmont, CA: Wadsworth.

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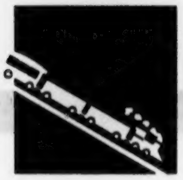
Appendix



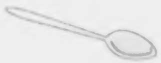
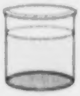
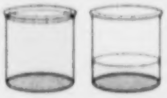
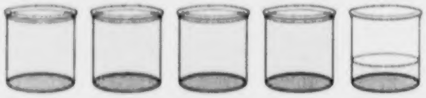
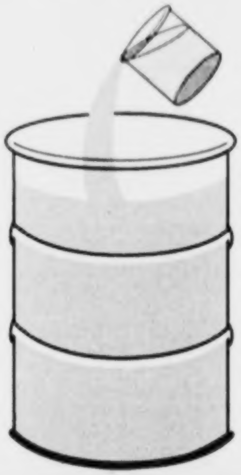
- **Transparency Masters**
  1. The Properties of Water
  2. Water Supply Distribution
  3. Dissolved Solids in Seawater
  4. Word Search
- **Humor**

The Properties of Water

Property	Compared to other substances
• Quantity present on earth's surface.	• Three times as abundant as all other substances combined.
• Dissolving ability	• Dissolves more substances in greater quantities than any other common liquid.
• Surface tension	• Highest of all common liquids.
• Conduction of heat	• Highest of all common liquids, except mercury.
• Heat capacity—quantity of heat required to raise temperature of 1 g of substance 1 °C.	• Highest of all common solids and liquids.
• Latent heat of fusion—quantity of heat gained or lost per unit mass by substance changing from solid to liquid or liquid to solid phase without accompanying temperature rise.	• Highest of all common solids and liquids.
• Latent heat of vaporization—quantity of heat gained or lost per unit mass by substance changing from liquid to gas or gas to liquid phase without an accompanying temperature rise.	• Highest of all common substances.
• Transparency	• Relatively great for visible light.
• Sound transmission	• Transmits sound well, compared to other liquids.

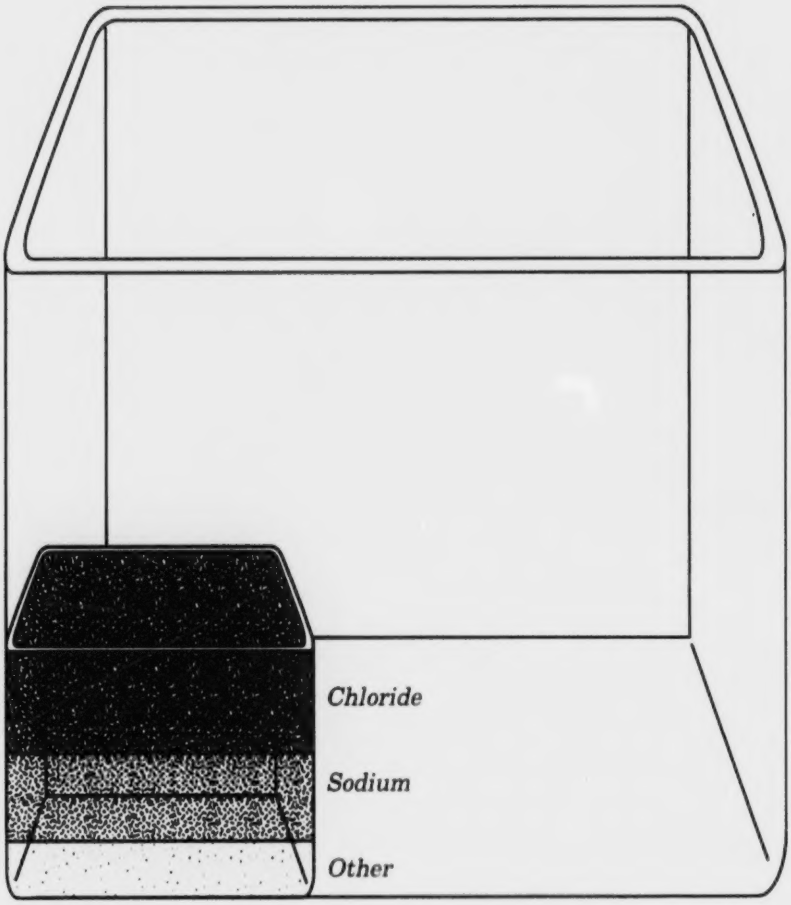
Water Supply Distribution

What if the world's total water supply were in a 200-L (about 55 gal) drum, then...

Source	***Representation***	***Percentage***
Rivers		0.0001
Atmospheric moisture		0.001
Lakes		0.016
Groundwater		0.61
Glaciers and ice caps		2.24
Ocean water		97.1

Adapated from ChemCom: Chemistry in the Community. (1988). Dubuque, IA: American Chemical Society, Kendall / Hunt Publishing Co.

Dissolved Solids in Seawater

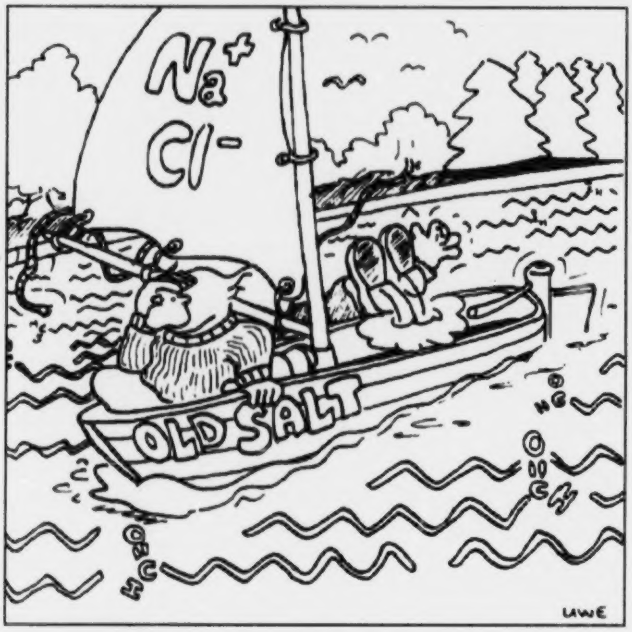

Chloride
Sodium
Other

Word Search

B	C	Y	U	K	X	R	Y	T	I	N	I	L	A	S	W	T
O	U	V	N	T	Y	X	Z	Z	N	Z	N	E	I	Q	B	N
I	Q	A	G	L	E	K	E	E	T	T	L	K	M	Y	E	I
L	C	C	N	S	C	V	P	C	G	E	L	Y	G	D	V	O
I	H	J	F	U	W	U	I	F	C	P	U	V	E	K	O	P
N	L	T	T	G	J	T	Q	T	X	G	G	X	Q	S	H	G
G	O	M	B	H	O	F	R	L	A	G	B	U	R	I	W	N
P	R	I	B	M	B	O	T	I	L	G	C	P	Q	S	H	I
O	I	W	S	R	L	L	C	S	U	G	I	M	P	O	N	Z
I	N	O	K	Y	S	Y	J	Z	M	C	B	L	X	M	W	E
N	I	P	T	U	I	N	L	B	Y	B	P	G	L	S	D	E
T	T	E	E	K	P	P	V	K	N	D	F	D	J	O	K	R
T	Y	T	Z	V	G	M	R	E	F	F	U	B	A	Z	C	F

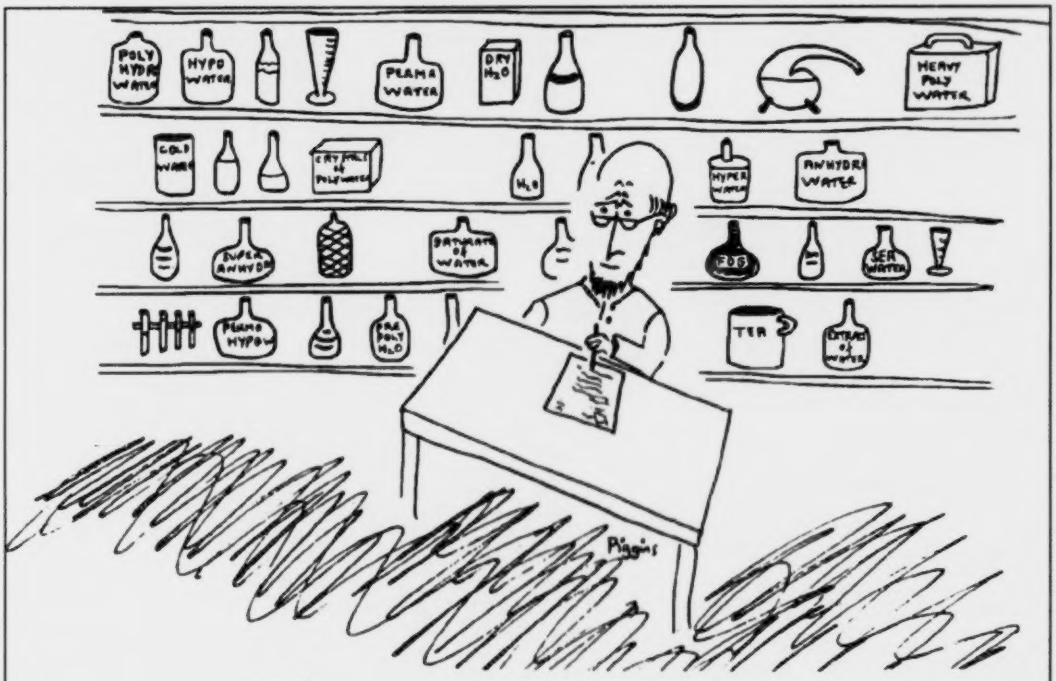
Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Total concentration of dissolved material (all ions), usually expressed in mol/L.
2. Flow of solvent into a solution through a semi-permeable membrane.
3. Total amount of dissolved halide ions present in a solution such as seawater.
4. Solution that resists change in pH when either OH^- or H^+ is added.
5. Adjective for physical properties of a solution that are different from those of a pure solvent and depending on the number of particles dissolved.
6. Pressure that must be applied to a solution to stop osmosis.
7. Temperature where the solid and liquid phase of a substance coexist (2 words).
8. Temperature at which the vapor pressure of the liquid is equal to the external pressure (2 words).
9. Unit of concentration abbreviation meaning one part of solute per billion parts of solution.
10. Substance whose solution in water conducts electricity.


Na+
Cl-
OLD SALT
UWE

Saline on Decanal

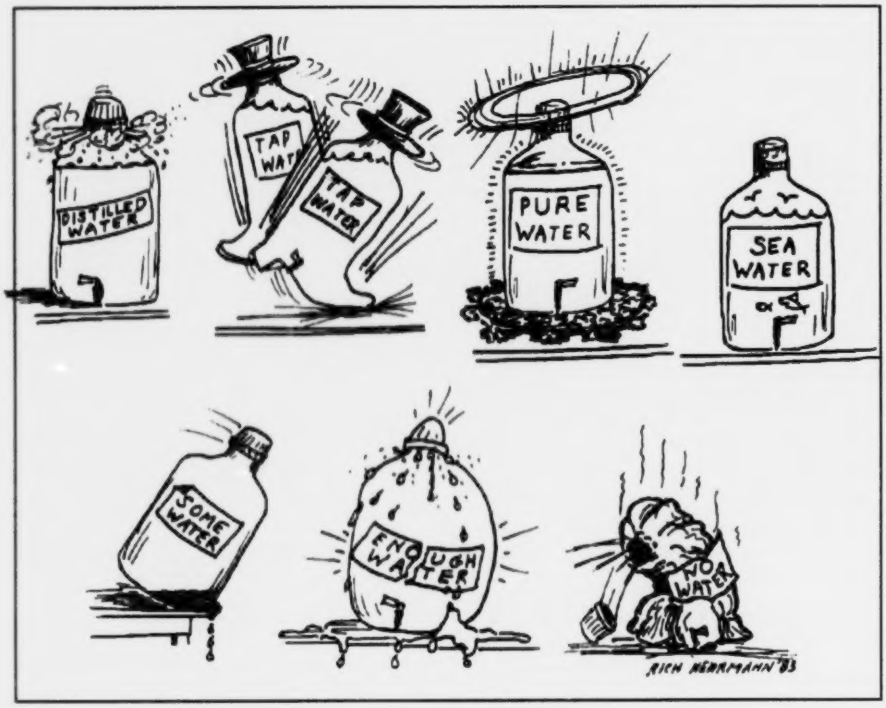
CHEM 13 NEWS, September 1979, p. 3. Reprinted with permission.


POLY HYDR WATER
HYPO WATER
PLAMA WATER
DRY H₂O
HEAVY POLY WATER
COLD WATER
CRYSTALS OF POLYWATER
H₂O
HYPER WATER
ANHYDR WATER
SUPER ANHYDR
SATURATE OF WATER
FOG
SEA WATER
PLAMO HYPO
PRE POLY H₂O
TEA
EXTRACT OF WATER
Piggins

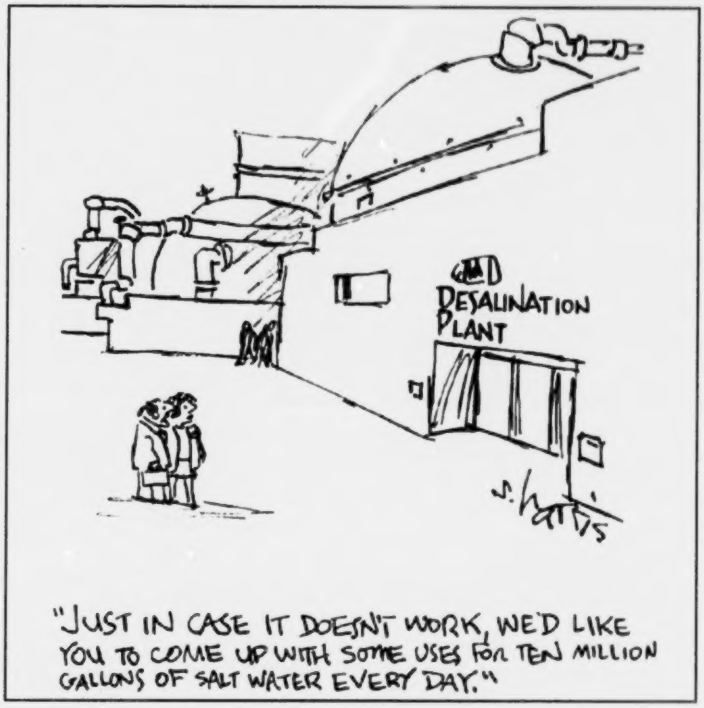
Dear Sir:

I understand that you have recently published an account of a new type of water ....

CHEM 13 NEWS, November 1969, p. 85. Reprinted with permission.


DISTILLED WATER
TAP WATER
TAP WATER
PURE WATER
SEA WATER
SOME WATER
ENOUGH WATER
NO WATER
RICH HERRMANN '83

CHEM 13 NEWS, April 1983, p. 1. Reprinted with permission.


DESALINATION PLANT
S. Harris

"JUST IN CASE IT DOESN'T WORK, WE'D LIKE YOU TO COME UP WITH SOME USES FOR TEN MILLION GALLONS OF SALT WATER EVERY DAY."

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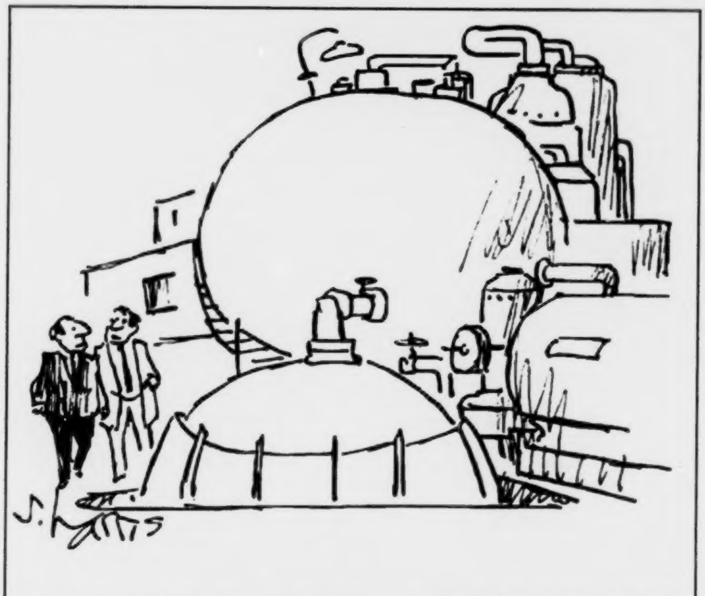

"OUR PLAN IS TO EXTRACT SULPHATES, BROMIDES, COPPER, SILVER AND GOLD FROM SEA WATER. ALL WE'VE MANAGED TO GET SO FAR, HOWEVER, IS SALT."

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A SourceBook Module

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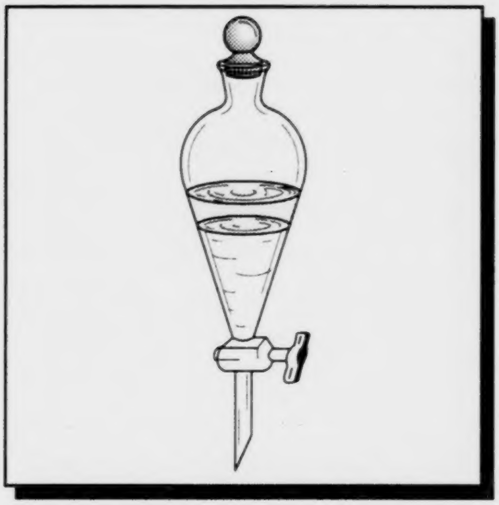
SEPARATIONS



ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Very few compounds occur pure in nature. They must first be mined or extracted and then *purified*. If a particular substance cannot be mined or extracted, then it must be *synthesized*. In any case, it is important to be able to *separate* desired compounds from others that are usually present. Although it does not sound exciting or seem glamorous, it is nonetheless, important in the modern practice of chemistry to be able to obtain pure compounds from the mixtures in which they occur. Four methods of separating the components of mixtures will be developed through laboratory activities. These are *filtration, fractional crystallization, fractional distillation, and chromatography.*

Separations involve the application of chemical principles, properties, and techniques to separate specific elements and compounds from mixtures. Such separations take advantage of differences in solubility, volatility, adsorbability, extractability, stereochemistry of isomers, and the properties of ions, elements, and molecules.

The mixtures to be separated may be either *heterogeneous* or *homogeneous*; in either case, suitable techniques are available to effect a satisfactory separation. Until fairly recently, many substances were scarce because suitable separation techniques had not been developed. For example, until the development of suitable chromatographic separation methods, the similar chemical properties of rare-earth (lanthanide) oxides did not allow a suitable separation. Now these oxides are available in quantity for a variety of uses.

PLACE IN THE CURRICULUM

This module probably does not stand alone in most chemistry courses; it may be partitioned into pieces to fit a particular need. For example, when considering physical properties, it might be well to include separations because the ability to separate depends upon differences in physical properties. Parts of the module may also fit into sections involving qualitative analysis, forensic chemistry, solutions, gases, mole, chemistry in medicine, and so forth. The possibilities are almost limitless.

CENTRAL CONCEPTS

1. Separation of two or more components of a mixture depends upon differences in physical properties. To effect such separations, a knowledge of the physical properties of the substances in a mixture is essential.
2. Filtration may be used for separation when one of two substances is insoluble in a given solvent by dissolving the soluble substance and filtering. The soluble substance passes through the filter but the insoluble one does not, *e.g.*, a mixture of sugar and powdered sulfur may be mixed with water, dissolving the sugar, which then can be passed through a filter with the sulfur remaining behind.
3. Fractional crystallization may be used for the separation of two substances with fairly large differences in solubility at a given temperature, *e.g.*, $NaClO_3$ (sodium chlorate) and NaCl (sodium chloride) at 0 °C have quite different solubilities allowing most of the $NaClO_3$ to crystallize with the NaCl remaining in solution.
4. Distillation may be used to separate a mixture of volatile and nonvolatile substances, *e.g.*, a salt water solution.

5. Fractional distillation utilizes the difference in boiling points of liquids, although some simple binary (two-component) mixtures form constant boiling azeotropes. For example, a mixture of 2-propanol and water may be separated into alcohol- and water-rich fractions.
6. Paper and column chromatography separate mixtures through differences in solubility in the developing solvent and differences in adsorptivity between the substances in the mixture and the developing substrate. For example, a mixture of food dyes may be separated using paper chromatography, or a mixture of chlorophylls in a green leaf may be separated using column chromatography.

RELATED CONCEPTS

1. Solutions are homogeneous mixtures with varying degrees of interaction between solute and solvent depending upon their nature, *i.e.*, polar or nonpolar.
2. Solubilities of substances is quite variable and depends, among other factors, upon the degree of interaction between solvent and solute. These differences provide a means of separation of substances.
3. Adsorption is a surface phenomenon that is useful for effecting separations.
4. Volatile substances vaporize easily while nonvolatile ones do not.
5. Melting and boiling points give an indication of the degree of interaction between the molecules and/or ions of substances.

RELATED SKILLS

Manipulative laboratory skills involving the use of common glassware and hardware are needed to perform the activities.

PERFORMANCE OBJECTIVES

After completing their study of separations, students should be able to:

1. describe the separation techniques used to isolate the components of mixtures and how to measure the amounts of the components. The techniques include but are not limited to: filtration, fractional crystallization, simple distillation, fractional distillation, and chromatography.
2. perform simple separations in the laboratory including filtration, distillation, crystallization and fractional crystallization, and chromatographic techniques (paper, thin layer, column, *etc.*).
3. describe some problem-solving applications of separations.
4. explain the importance of separation of components of mixtures to the maintenance of our quality of life and leisure time, particularly the importance of the separation techniques in industry and production of the commodities of commerce.
5. describe some of the applications of separation techniques, *e.g.*, quality control, consumer protection, pollution monitoring.
6. describe some uses of separations in their communities, *e.g.*, water quality testing, sewage treatment, medical applications, *etc.*
7. list some of the vocational and career opportunities available that are based on separation techniques.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Separating Mixtures

Introduction

To determine the chemical and physical properties of elements and compounds, chemists need to be sure that the substances are pure. However, most substances in nature are mixtures and must be separated into their components. To effect such separations, it is possible to use differences in properties by employing appropriate techniques. The techniques presented in this activity are fairly representative of those that have proven useful in both the laboratory and in industry for this purpose. Additional techniques may be demonstrated.

Purpose

To illustrate several commonly used techniques for separating the components of a mixture.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. The usual laboratory safety precautions apply to the four procedures presented in this activity. However, particular caution should be exercised with some of the substances used since these may be toxic or flammable.
3. Open flames should not be used when alcohols or other organic solvents are present.
4. Observe the safety rules presented prior to each procedure.

Part I. Gravity Filtration

Purpose

To separate soluble and insoluble substances by dissolving the soluble substance in water, and then separating the insoluble substance from the liquid solution by filtration.

Safety

Treat the mixture to be separated with the usual precaution, *i.e.*, do not taste or ingest the mixture. Dispose of solid in an appropriate container. Do not return excess mixture to its container.

Procedure

1. Obtain an unknown mixture from your teacher. Weigh out 3.00 g unknown to the nearest 0.01 g on a piece of weighing paper. It is not important that exactly three grams is weighed out, but it is important that the weight is known to the nearest ± 0.01 g. Transfer the weighed unknown to a 250-mL beaker. Add 100 mL distilled water and stir.
2. Weigh a filter paper circle to the nearest two decimal places and record the mass. Prepare the filter paper circle by folding it in half, and then again into quarters. Open the filter paper to form a cone with one thickness of paper on one side and three thicknesses on the other. Place the filter paper cone into a filter funnel, place the funnel in a ring on a ringstand, and then moisten the paper with a small volume of distilled

water from a wash bottle. Gently press the paper against the sides of the funnel to make a snug fit by using a glass stirring rod. If correctly selected, the filter paper cone's top edge will be just below the rim of the funnel.

3. Use a 250-mL beaker to collect the **filtrate** (the liquid component of the mixture) by placing the beaker below the funnel tube so that the tube touches the side of the beaker. The funnel tube should extend below the beaker's rim by 2-4 cm. After stirring, pour the mixture from the first beaker down a glass stirring rod into the funnel. Do not allow the liquid level in the funnel to be more than three-fourths of the distance from the bottom of the funnel to the top of the filter paper.
4. When filtration is complete, rinse any remaining solid particles from the first beaker into the funnel by using a wash bottle and distilled water. Pour three 10-mL portions of distilled water over the solid in the filter paper. Remove the beaker with filtrate and replace it with a clean beaker. Then pour three 10-ml portions of acetone over the solid in the filter paper to rinse it free of water. Remove the filter paper and contents. Carefully open the filter paper circle and place it on a paper towel, being careful to not lose any of the solid. Place the paper towel with filter paper and solid under the fume hood to air dry. Pour the rinse acetone from the beaker into the container designated by your teacher for this purpose. Dispose of the filtrate as directed by your teacher.
5. After the filter paper and solid are dry, weigh them to the nearest two decimal places on the laboratory balance. The difference between the mass of solid collected and the original sample allows the composition of the mixture to be determined.
6. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Using the difference in mass between the original unknown sample and the solid obtained after filtration, calculate the percent composition of the original sample as percent by mass of the soluble component of the mixture.
2. Obtain the accepted value from your teacher and calculate percent error as follows:

$$\% \text{ Error} = \frac{\text{Accepted value} - \text{Experimental value}}{\text{Accepted value}} \times 100\%$$

Implications and Applications

1. How efficient is gravity filtration? Explain on the basis of your data.
2. If the filtrate in your activity is not as clear as distilled water, offer a possible explanation.
3. Why should the end of the filter funnel touch the side of the collecting beaker in gravity filtration?

Part II. Fractional Distillation

Purpose

To separate a mixture of two miscible liquids (liquids that mix in all proportions) with different boiling points. Your mixture will consist of water and 2-propanol (isopropyl or rubbing alcohol).

Safety

1. Wear protective goggles throughout the laboratory activity.
2. 2-propanol (rubbing alcohol) is generally safe; however, it is flammable, toxic, and requires care in handling.
3. Use open flames with caution since the alcohol and some alcohol-water mixtures are flammable.
4. Place the alcohol-water mixtures in a designated container.

Procedure

Carefully examine the mixture that is to be separated. Cautiously smell the mixture and note any odor. Dip a piece of filter paper in the mixture to wet one end and try to light it with a match. Carefully measure the density of the mixture using a graduated cylinder and balance. Test the solubility of sugar in a small amount (about 1 mL) of the liquid. Record all observations.

1. Set up a distillation apparatus as shown in Figure 1. Check the following details.

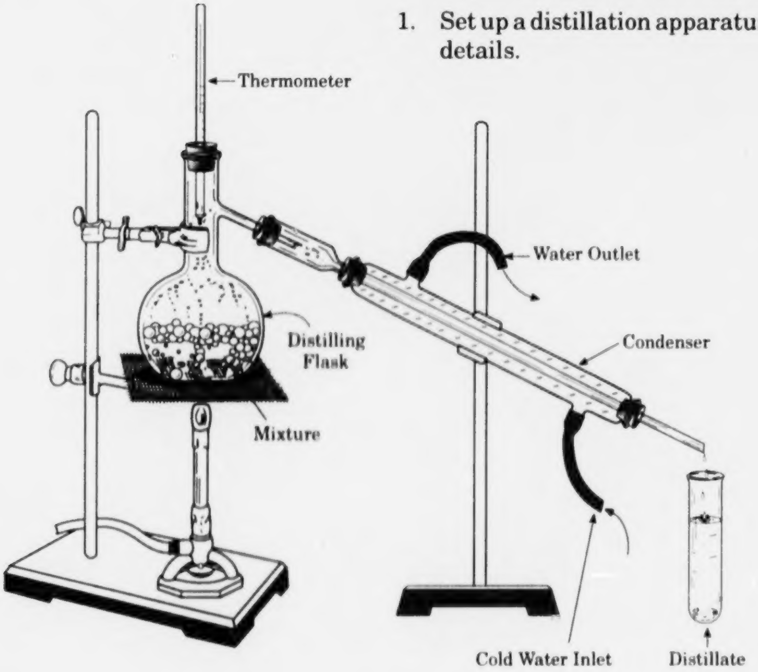


Figure 1. A simple distillation apparatus.

   a. The top of the thermometer bulb should be just below the side arm of the distilling flask, insuring that the entire thermometer bulb will be bathed in the rising vapor as the liquid mixture is heated.
   b. The side arm should extend beyond the cork in the top end of the condenser, and the condenser tip should extend beyond the cork in the adapter. The distillate can not pick up impurities from the corks when set up this way.
   c. Both the distilling flask and the condenser must be firmly clamped in place.
   d. To prevent having a closed system, the lower end of the adapter should not be connected to the receiver by a cork. This arrangement prevents pressure from building up and causing an accident.
   e. Begin circulating water from the lower end of the condenser prior to heating the distilling flask.
   f. Add two or three boiling chips (small marble chips or porcelain chips) to the distilling flask to prevent superheating and bumping.
   g. Have your teacher check the setup before applying heat to the flask.

2. Label and number four large test-tubes (20- x 150-mm) for collecting fractions. Place Tube 1 under the condenser outlet.

3. Remove the thermometer with its cork. By means of a long-stem glass funnel, add about 20-30 mL 2-propanol-water mixture to the distilling flask.
4. Gently heat the flask until the liquid begins to boil. Take temperature readings every half-minute and record these along with the time. Continue heating just enough to keep the liquid boiling while continuing to take temperature readings. Continue boiling the liquid until the temperature remains the same for about three or four minutes.
5. Graphing the data from Step 4 will provide the information needed to decide about the temperature range to collect separate fractions. Do the graphs before continuing. Repeat the procedure (Steps 3-4) but this time use the data collected in Step 4 to collect separate fractions of distillate in Tubes 2, 3, and 4. You should be able to detect the presence of at least three separate fractions, the last one being essentially pure water.
6. Cautiously smell each fraction and record your observations. Test the flammability of each fraction by dipping a small piece of paper in each and trying to ignite the wet paper with a match. Use a balance and graduated cylinder to measure the density of each fraction. Test the solubility of sugar in a small amount (about 1 mL) of each fraction. Record all observations.
7. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Can you tell just by looking that the clear liquid is a mixture?
2. Does the paper dipped in the liquid burn?
3. What is the density of the mixture prior to distillation?
4. What is the density of each fraction?
5. Does sugar dissolve in each fraction?

Implications and Applications

1. What do you conclude about the composition of each fraction?
2. On the basis of the tests performed on the original solution and each fraction, what can you conclude about the composition of each fraction?
3. What do you think would be the result if you redistilled each fraction into three fractions of equal volume?

Part III. Fractional Crystallization

Purpose

To separate a mixture of two solutes dissolved in the same solution by using differences in solubility at a low temperature.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Sodium chloride and potassium nitrate are generally safe to handle; however, potassium nitrate is an excellent oxidizer and *must not be in contact with organic substances or other oxidizable materials or metals.*

Procedure

1. Examine the solid sample supplied by your teacher. Is it possible to see more than one substance? Record observations.
2. Transfer the solid to a 150-mL beaker. Add about 10 mL distilled water to the mixture, and heat gently while stirring with a glass stirring rod.
3. When the solution boils, add more distilled water in small portions with heating and stirring until the solid just completely dissolves. About 25 mL water should be sufficient. *Do not add too much water or precipitation may not occur.*
4. Add two or three boiling chips, cover with a watch glass, and heat to boiling until about one-half of the water has evaporated. Some spattering may occur but will not affect the the results. While waiting for the water volume to reduce, prepare a funnel and filter paper. Obtain a test-tube large enough to hold the remaining liquid (at least 25- x 150-mm).
5. Just before doing the next step, pour about 15-20 mL hot distilled water through the filter paper and the funnel that has been prepared.
6. Immediately filter the hot solution into a clean, large test-tube. Use beaker tongs or a paper towel folded into a narrow strip encircling the beaker for handling the hot beaker. Scrape any solid remaining in the beaker onto the filter paper.
7. If instructed to do so, allow the filtrate (clear solution) to cool overnight and perform Steps 6-8 in the next laboratory period. Otherwise, place the test-tube containing the filtrate into an ice-water bath in a 400-mL beaker. Allow crystals to precipitate from the solution. When precipitation is complete, remove the test-tube from the water bath, decant the filtrate, scrape the solid out of the test-tube with a stirring rod, and dry the solid on a filter paper circle. Folding the filter paper, placing it between paper towels, and then squeezing expedites removing water from the residue.
8. The original filter paper circle contains mainly sodium chloride while the second filter paper contains mainly potassium nitrate. Compare the appearance of the two solids. Describe any similarity or difference between the two.
9. Place 0.5 g amounts of each solid in two separate large test-tubes and add 2 mL distilled water to each. Shake or stir to dissolve each solid. Then, carefully heat each test-tube to boiling and continue boiling gently until the solutions become saturated, *i.e.*, some solid precipitates out of solution. Qualitatively compare the solubilities of the two solids.
10. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Do the two solids look alike after they have been separated?
2. How do the solubilities of the two solids compare in cold and hot water?
3. Are the two solids two different substances?

Implications and Applications

1. What can you conclude about whether the two solids have been separated?
2. If 100 mL water at 10 °C is saturated in both potassium nitrate and sodium chloride, predict what will happen if the temperature of the solution is raised to 100 °C.

3. The solubility of both sodium chloride and potassium nitrate is about 36 g/100 g water at 23 °C, but at 100 °C, the solubilities are about 39 g/100 g water and about 240 g/100 g water, respectively. At 10 °C, the solubilities are about 36 g/100 g water and about 21 g/100 g water, respectively. Why is the solubility difference of these two substances at the temperatures given advantageous in separating the substances?

Part IV. Chromatography

Purpose

To separate the colored substances in an extract from green leaves by column chromatography.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. The solvents used are highly flammable; open flames must not be used.
3. Do not pour solvents into the drain but dispose of them as directed by your teacher.

Procedure

1. To prepare a chromatography column, obtain a 25- x 2-cm glass tube, a one hole rubber stopper fitted with a glass tube, a short length of latex rubber tubing to fit the glass tube, a screw (or pinch) clamp, a plug of glass wool, and a short medicine dropper. A slurry made of 50 g of washing soda (sodium carbonate) and TTE (1,1,2-trichlorotrifluoroethane) will be used to pack the column. Assemble the chromatography column as shown in Figure 2.
2. Using a ring stand and support clamp, clamp the tube in a vertical position. After the screw (or pinch) clamp is closed, pour a slurry made from 50 g washing soda mixed with just enough TTE to be pourable into the upright tube. Use a thin glass rod to pack the slurry in the tube. It is important that there are no channels between the slurry and the tube. The column of washing soda should be about 20-cm tall. Run excess TTE from the bottom of the tube until the meniscus is just at the surface of the solid. If the column must be stored, it should be stoppered to prevent solvent evaporation.
3. Prepare three clean, dry test-tubes for collecting the fractions eluted. Label the test-tubes: xanthophyll and carotenes, Chlorophyll A, and Chlorophyll B.

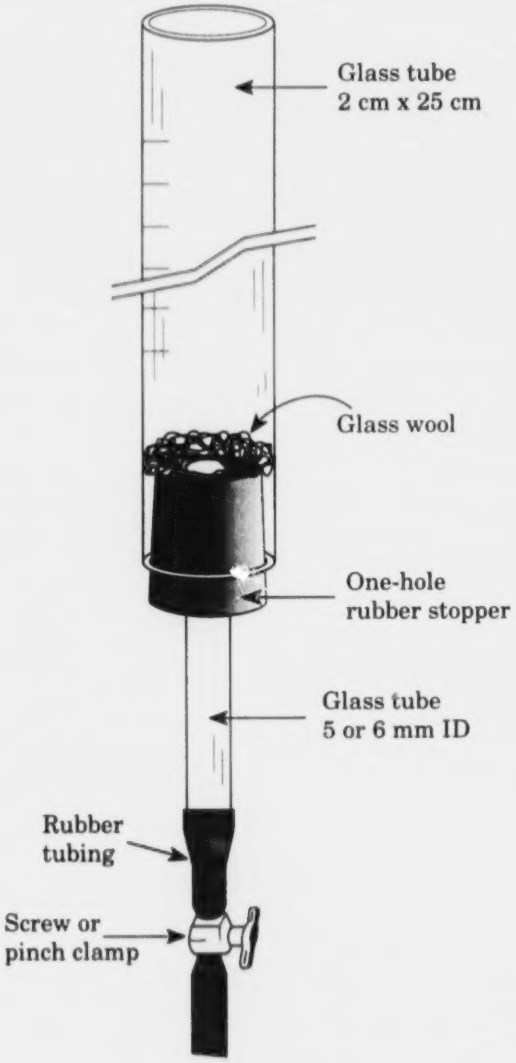


Figure 2. Column chromatography apparatus.

4. Obtain 20 mL green plant extract and *carefully* pour it into the column. Try not to disturb the top of the column. Use a 125-mL conical flask to collect the **eluate** (liquid running from the column), and adjust the clamp for a flow rate of one drop per second.
5. When all of the sample has been adsorbed onto the column, add 10 mL TTE to the top of the column. Two bands of xanthophyll and carotenes should separate from the dark green band at the top and move rapidly down the column. Collect this yellow fraction (about 6 mL) in the test-tube. When the TTE layer is nearly at the top of the column, 10 mL of 10% solution of acetone in TTE is added. A band of chlorophyll A (yellow-green) should separate and move down the column. Collect this fraction (about 7 mL) in a separate, labeled test-tube.
6. After the Chlorophyll A is collected, add 10 mL of 20% acetone in TTE solution to remove Chlorophyll B. Collect this dark green sample in a separate, labeled test-tube.
7. Follow your teacher's directions for cleaning the column.
8. *OPTIONAL:* If directed to do so, confirm the presence of the two chlorophylls by taking an absorption spectrum with a spectrophotometer. Each should show a peak absorption at a characteristic wavelength (see *Instrumentation* module).
9. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. How many *different* color bands were observed as the sample of plant pigments traveled through the chromatography column?
2. Was the separation complete or did the bands seem to run together? What might cause the bands to run together?
3. *OPTIONAL:* What was the wavelength of maximum absorption for the two green colored substances when the substances' absorption spectra were taken with a spectrophotometer?

Implications and Applications

1. Why were two different concentrations of acetone in TTE used?
2. Why did the mixture of pigments separate? Consider both solubility and adsorption by the substance in the chromatography tube.
3. It is not unusual to find that some substances that frequently occur together have similar chemical and physical properties, *e.g.*, compounds of the lanthanide and actinide elements. Could chromatography be useful to separate such substances?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Separating Mixtures

Major Chemical Concept

Instrumental analysis is commonplace in the practice of chemistry. However, it is necessary to have relatively pure samples of substances to analyze. Producing such samples is the role of physical separations since sources of most compounds are rarely pure. After the substance is isolated and purified, instrumentation may be used to study its identity and structure.

Separation procedures presented in this activity are relatively simple and all rely upon a differential physical property. Although procedures presented here are all single step operations, they can be linked to form multi-step operations.

Some mixtures of solids may be separated by dissolving one or more components. For example, a mixture of sand, salt, and iron filings is such a mixture. This mixture is classified as **heterogeneous**. This means it is not uniform but varies in composition from point to point. This fact would be obvious under a microscope. It would be possible to separate the three components of the mixture by using a microscope, tweezers, and a large amount of time and patience. A much more rapid separation is possible, however, by using other methods. A magnet can be used to remove the iron from the mixture leaving the sand and salt. Water can then be added to dissolve the salt, and the resulting solution separated from the sand by **filtration**. The solution passes through the porous filter while the sand does not. Finally, the water may be evaporated to leave the salt residue. This evaporation could be a simple **distillation** that is done by heating the solution in a flask and passing the vaporized water through a condenser. In this fashion the added water may also be collected in fairly pure form.

Distillation can also be used to separate volatile liquids by a process known as **fractional distillation**. In this procedure, the liquid mixture is placed in the flask of a distillation apparatus. As the flask is heated, the lower boiling liquid vaporizes faster than the higher boiling liquid. Thus, although both components are in the vapor phase, the vapor is richer in the lower boiling liquid. The resulting vapor could be condensed and re-evaporated making the resulting vapor still richer in low boiling liquid. If the process is repeated a sufficient number of times, a separation of the two volatile components can be achieved.

Crystallization, or **fractional crystallization**, is a purification method that depends upon the change of solubility of a substance with a temperature change. In practice, when two substances have distinctly different solubilities, this method may be used. For example, mixtures of KNO_3/NaCl and $NaClO_3$/NaCl can be separated this way. The process is based upon the fact that solubilities of KNO_3 and $NaClO_3$ are much greater than that of NaCl in hot water, but much lower in ice water. If a mixture of either pair is dissolved in hot water, and then the solution is allowed to cool to 0 °C, the more soluble $NaClO_3$ will crystallize out of solution while little if any of the NaCl crystallizes, provided that a sufficiently small amount of NaCl was initially present. To further purify the material, it may be recrystallized. A complete separation of the two substances is not made; however, a large proportion of the more soluble substance is recovered in rather pure form. The remaining solution will contain some of both compounds.

Chromatographic methods may be used to separate liquids and solids. Chromatography constitutes one of the most versatile and powerful separation techniques used by chemists. The types of chromatography considered here are paper chromatography and column chromatography or liquid-solid chromatography.

Other common chromatography types are thin layer chromatography (see *Forensic Chemistry* module) and gas-liquid chromatography. **Paper chromatography** depends upon a piece of hanging paper with one end dipping in a liquid so that the liquid wicks up along the paper. A drop of solution containing a mixture to be separated is placed near the edge of the paper strip dipping in the liquid. When the liquid in contact with the paper wicks upward, it passes over the drop of mixture. The components in the drop that interact least strongly with the paper and most strongly with the liquid travel farther up the paper with the wicking liquid. In this way, components of the mixture are separated.

Level

Procedures I, II, and III may be used with all levels of students. Procedure IV is better used with honor students, and may be demonstrated to students at other levels. The choice of procedure(s) will depend upon your confidence, apparatus and materials available, and students' ability level.

Expected Student Background

Students should have some experience with laboratory apparatus and glassware. These procedures require good manipulative ability. You may need to demonstrate some of the techniques such as filtration and setting up the apparatus. Students should also know about properties such as melting point, solubility, adsorption, *etc.*, since these properties form the basis for the separations in the activity.

Time

Teacher preparation time for each procedure varies between 30 min to 2 hr, depending upon how accessible the equipment is and how much of the equipment you plan to assemble for students. Chemical preparation is at a minimum with either simple solutions to prepare or grinding a mixture of NaCl and KNO_3 together. Students should be able to finish the procedures for simple filtration and fractional crystallization in a single period. The procedures for fractional distillation and column chromatography will each require two laboratory periods (50-55 min).

Safety

Read the *Safety Considerations* in the *Student Version*. Treat all the substances used in these procedures as potentially hazardous. Instruct students about disposal of the solvents. 2-propanol can be safely flushed down the drain with sufficient water. TTE (1,1,2-trichlorotrifluoroethane) should be collected and recycled. All solids may be disposed of in the usual fashion. It is recommended that you collect the potassium nitrate since students might be tempted to take some for use in making gunpowder. Safety goggles must be worn in the laboratory at all times. Instruct students to wash their hands thoroughly prior to leaving the laboratory.

Materials (For 24 students working in pairs)

Part I. Simple Filtration Procedure

Nonconsumables

Balance, centigram
12 Wash bottles
24 Beakers, 250-mL
Distilled water
12 Graduated cylinders, 100-mL and 10-mL
12 Filtering funnels
12 Glass rods, stirring
12 Spatulas, porcelain

Consumables

- 12 Filter paper circles
- 50-50 Mixture NaCl and $CaCO_3$, 36-48 g
- Weighing paper
- Paper towels

Part II. Fractional Distillation Procedure

Nonconsumables

- 12 Distilling flasks, 150-mL
- Boiling chips
- 12 Condensers fitted with rubber tubing
- 24 Ring stands with support clamps
- 12 Iron rings
- 48 Test-tubes, small
- 12 Stirring rods
- 12 Wire gauze, ceramic centered
- 12 Burners
- 12 Thermometers, –10 °C to 110 °C x 1 °C
- 12 Adapters to fit condenser (not absolutely essential)
- 60 Test-tubes, 20- x 150-mm
- 12 Funnels, long stem
- 12 Graduated cylinders, 10-mL, 25-mL, and 50-mL
- 12 Test-tube racks

Consumables

- Sugar, 10 g
- Boiling chips
- Filter paper or other porous paper
- 2-propanol-distilled-water mixture, 1200 mL (two parts 2-propanol to three parts water)
- Matches
- Labels
- Marking pencils
- Graph Paper
- Rulers
- Pencils

Part III. Fractional Crystallization Procedure

Nonconsumables

- 24 Beakers, 150-mL and 250-mL
- 12 Beakers, 400-mL (optional for Step 6 ice-water bath)
- Boiling chips
- 12 Filtering funnels
- 12 Graduated cylinders, 10-mL
- 12 Watch glasses (to cover beaker)
- 12 Stirring rods
- 36 Test-tubes, 25- x 150-mm

Consumables

- 24 Filter paper circles to fit funnel
- 50-50 Mixture of NaCl and KNO_3, 200 g
- Distilled water
- Boiling chips

Part IV. Chromatography Procedure

Nonconsumables

- 24 Beakers, 250-mL
- 12 Conical flasks (Erlenmeyer), 125-mL
- 12 Glass tubes, 25- x 1- or 2-cm (2 cm diameter is preferred) or 12 Mohr burets, 25- or 50-mL
- 12 Rubber stoppers, one-hole to fit glass tube
- 12 Rubber stoppers, solid to fit glass tube
- 24 Graduated cylinders, 10-mL
- 12 Short glass tubes to fit one-hole rubber stopper
- 12 Latex rubber tubes, 4-in length
- 12 Clamps, screw or pinch to fit rubber tubing
- 12 Short medicine dropper glass tubes
- 12 Ringstands with support clamps
- 36 Test-tubes
- 12 Stirring rods

Consumables

- Sodium carbonate, Na_2CO_3, 600 g
- TTE (1,1,2-Trichlorotrifluoroethane), 2.4 L
- Glass wool (small amount for chromatography tube)
- Plant extract (see *Advance Preparation*)

Advance Preparation

One activity is presented for use with this topic. It is about separations and provides four procedures including simple filtration, fractional crystallization, fractional distillation, and column chromatography. Any number may be used depending upon your objectives. Another possibility is to divide the class into four groups, have each group perform one of the procedures, and then give a class report. Check to be certain that the safety shower, the fire blanket, the eyewash, and the fire extinguisher are functioning correctly.

Part I

For the simple filtration procedure, prepare a 50-50 mixture of precipitated chalk ($CaCO_3$) and sodium chloride. (If precipitated chalk is unavailable, grind white chalk to a fine powder in a mortar and pestle.) Allow about 5-6 g per student or group. Grind the sodium chloride in a mortar and pestle until it is powdered. Mix the two substances thoroughly so that the mixture appears to be a single substance.

Part II

Prepare a 2:3 mixture (by volume) of 2-propanol and distilled water. Ordinary rubbing alcohol may be used; however, the rubbing alcohol is a 70% by volume mixture of 2-propanol and water. Thus, approximately equal volumes of the alcohol and water will be about right.

If condensers and distilling flasks are unavailable, a large test-tube fitted with a two-hole rubber stopper having a thermometer and a bent glass tube can be used. The thermometer bulb should be close to the bottom of the stopper and the bent glass tube is fitted with clear plastic tubing that leads into a large test-tube immersed in a beaker of cold water. The water-alcohol mixture is heated in the test-tube in this alternate setup. Of course, a smaller amount of the mixture to distill will be needed, *i.e.*, only about 40 mL per student or group.

Part III

The unknown mixture is prepared by grinding together in a large mortar and pestle equal masses of reagent grade sodium chloride (do not use common table salt) and potassium nitrate. For safety, make certain that no combustible material is in the mortar prior to grinding. Each group or student should be given about 15 g of the mixture. To simplify this task, place a 15-g sample in a test-tube, and then fill the other test-tubes to the same level. This will save considerable time rather than having each student or group weigh out a sample. Using 100 g of each substance in the mixture is adequate for a class of 24 students working in pairs.

Since the solution must be filtered while hot, the funnel and filter paper need to be preheated with hot distilled water immediately before the solution is filtered. A second burner, ring stand, iron ring, and ceramic centered wire gauze is needed for this purpose, or a large container of hot distilled water can be made available for students to obtain hot water.

Part IV

To prepare the extract, obtain about 10 g of a green plant (frozen spinach works well). Homogenize the green plant with 25 mL acetone in a blender for 1-2 min. (Alternately, the green plant may be ground in a mortar and pestle with some clean sand, and the green pigments extracted with 25-35 mL acetone. The remainder of the procedure is the same after either method of extraction.) The homogenate is then suction filtered (Buchner funnel with fritted glass or filter paper circle to fit), suction flask, tubing, and aspirator. The residue in the funnel is washed with 10 mL acetone. Discard the residue. The filtrate should be a very dark green solution. Extract the filtrate with 30 mL TTE in a large separatory funnel.

After extraction, discard the acetone portion and wash the TTE portion (green color) three times with 10 mL distilled water in the separatory funnel. Discard the water, place the TTE solution in a flask along with 5 g anhydrous Na_2SO_4 to absorb the residual water. Stopper the flask and set aside in a dark place. To save time, the chromatography tubes may be prepared in advance, stoppered, and stored until needed. While in storage, the TTE level in the chromatography tubes should be just above the sodium carbonate in the tube. *NOTE*: In preparing the chromatography column, be sure to make a fine slurry of sodium carbonate and TTE. It should pour easily into the column and be allowed to settle to avoid channeling. Some teachers prefer to use powdered sugar as the stationary phase and hexane as the mobile phase.

Pre-Laboratory Discussion

Select one or two procedures that are within student capabilities. Thoroughly discuss the procedure and demonstrate any new techniques prior to students' laboratory work. Caution students about the flammability of solvents used and the danger of having open flames while solvents are being used. Another possibility is to have different student groups do different procedures and make a group report about the procedure and results. However, do not allow procedures using acetone to be performed next to procedures requiring a burner. Hotplates can be substituted as a source of heat.

Help students with data tables and/or methods of recording observations. None of the procedures require elaborate data tables, so designing them should be within the capability of the student. Point out that measurements must be recorded. It might be sufficient to simply identify the data to be recorded and leave the rest to the student.

Equipment and materials for the laboratory should be easily accessible to students. Demonstrate the assembly of apparatus for *Parts II* and *IV*. For *Part II*, demonstrate the distillation assembly consisting of distilling flask, condenser, and adapter. Include the direction of water flow through the condenser (bottom to top—see Figure 1). Most chemistry texts have solubility graphs showing the solubility of NaCl and KNO_3. You may want to include this in the pre- or post-laboratory discussion for *Part III*.

Teacher-Student Interaction

Circulate to answer student questions and help with equipment setup and technique. Acting as a safety officer is appropriate. Do not hesitate to correct students for unsafe practices. Encourage students to complete assigned procedures and report results regardless of the student's view about correctness of results. If students decide better results can be achieved by repeating the procedure, encourage them to do so if there is time. Providing students with specific uses for separation techniques should be motivating.

Anticipated Student Results

Part I. Gravity Filtration

With reasonable care, students should obtain results within 5-10% of the actual composition of the original mixture. The method of calculating the composition is given:

$$\frac{\text{Mass of mixture} - \text{Mass of residue}}{\text{Mass of mixture}} \times 100 = \%\ \text{Soluble component}$$

Students should note that filtration of the mixture is a relatively slow process. Some students will note that the filtrate is not clear and colorless like distilled water. If the filtrate is not clear, consider refiltering to clarify the filtrate. The quality of the filter paper may dictate whether the filtrate is clear or not.

Part II. Fractional Distillation

Students will probably agree that it is impossible to tell that the original solution is a mixture of liquids. If heating is done slowly and carefully, students should obtain three separate fractions of liquid. Fraction 1 appears to be a pure substance having a boiling point of about 81 °C; Fraction 2 is a mixture of substances similar to the original solution; and Fraction 3 seems to be a pure substance with a boiling point the same as water. (Actually, if properly collected, Fraction 1 is a constant boiling azeotrope that is 90% alcohol and 10% water with a constant boiling point of 80.5 °C. It cannot be further separated by distillation, and thus it has the appearance of a single substance. Do not dwell on the purity of this fraction but emphasize that a separation has occurred.) The initial heating and subsequent graph of temperature *vs.* time is the key to determining the temperature ranges over which to collect the separate fractions (see Figure 3 for "typical" data).

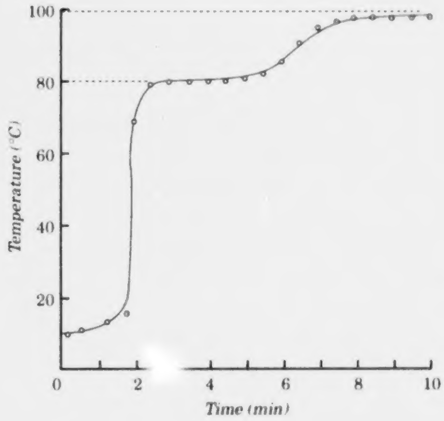


Figure 3. Typical student data for fractional distillation activity.

There are two distinct plateaus or flat segments indicating that there are two different substances in the mixture, leading to the suggestion that three fractions should be collected to separate the two substances. Fraction 1 should be collected about 78-82 °C, Fraction 2 from 82 °C to about the boiling temperature of water, and fraction 3 beginning at about the boiling point of water. (The boiling point of both liquids depends upon atmospheric pressure at time of the activity. Boiling points are lowered by a decrease in pressure.)

A typical set of properties determined for the solution and three fractions is shown in the following table.

Property	Solution	Fraction 1	Fraction 2	Fraction 3
Flammability	Yes	Yes	Yes	No
Odor	Medium	Strong	Medium	None
Density	~0.91 g/mL	~0.084 g/mL	~0.93 g/mL	~1.00 g/mL
Boiling point	Variable	~81 °C	Variable	100 °C
Dissolved sugar?	Yes	No	Slight	Yes

In order to obtain density data comparable to that listed above, students will need to make the mass and volume measurements carefully.

Part IV. Chromatography Procedure.

Students may not see actual separation of fractions on the column, but elution will allow them to see clearly three fractions, a yellow fraction absorbing at about 430 nm, a yellow-green fraction with absorption maxima at 460 and 600 nm and a green fraction with absorption maxima at 480 and 650 nm.

Answers to Implications and Applications

Part I. Gravity Filtration

1. Students will answer that gravity filtration is efficient at separating soluble and insoluble substances, but that it is not efficient from the standpoint of speed.
2. If the filtrate is unclear, some of the insoluble chalk passed through the filter paper. This is due to overfilling the funnel (so that some chalk slurry runs down between the filter paper and funnel), the filter paper is too porous to retain the smallest chalk particles, or to a hole punched in the filter paper.
3. By having the end of the funnel touch the side of the beaker, an unbroken liquid column is provided. This facilitates the flow and prevents spattering by having drops strike the liquid surface.

Part II. Fractional Distillation

1. Students should be able to conclude that a separation of two liquids has taken place on the basis of the evidence gathered by testing flammability, odor, and density of collected fractions.
2. On the basis of density, flammability, and odor tests, the first fraction was almost pure alcohol; the second fraction was quite similar to the original solution, and the third fraction was water.
3. If the fractions were redistilled, the following are the predicted results. Fraction 1 would boil at about 80 °C until it was all gone. Fraction 2 would come to a boil at about 80 °C and after all the alcohol was gone, the temperature would rise to the boiling point of water and remain there until all the liquid boiled away. Fraction 3 would begin to boil at about the boiling temperature of water and remain at that temperature as long as liquid remained.

Part III. Fractional Crystallization

1. The appearance of two solids and differences in solubility allow the conclusion that a separation has occurred.
2. Since the solutions are saturated at 10 °C, heating to 100 °C would not produce an observable change other than dissolving some of the undissolved solute present.
3. The sodium chloride and potassium nitrate may be separated by fractional crystallization because of the large difference in solubility at 100 °C and because the solubility of sodium chloride is nearly constant over the temperature range of 10 °C to 100 °C. When the hot solution is filtered, it contains almost pure potassium nitrate. Any excess sodium chloride remains in crystalline form. The filtrate, then, has little sodium chloride present, and when allowed to crystallize, the resulting solid is almost entirely potassium nitrate. If the solid is then put into solution and allowed to crystallize again, the residual sodium chloride remains in solution. Filtering and washing with water will result in removal of most of the sodium chloride.

Part IV. Column Chromatography

1. TTE removes rapidly the xanthophyll and carotenes. Chlorophyll and sodium carbonate strongly interact. Adding acetone, a polar solvent, reduces the interaction between chlorophyll and sodium carbonate by decreasing the number of adsorptive sites on the stationary phase (sodium carbonate). Chlorophyll A interacts with sodium carbonate less than does Chlorophyll B; hence, a smaller concentration of acetone is needed to help it move down the column.
2. A mixture like that in this activity may be separated in a chromatographic column because of different degrees of interaction of the pigments with sodium carbonate and the solvent.
3. Chromatographic separations are successful in many cases where other techniques fail due to the similarity in properties of the substances to be separated. The actinide oxides, for example, may be successfully separated in this way. These oxides could then be converted to the actinide element.

Post-Laboratory Discussion

Students will require minimal help with the data analysis for the simple filtration procedure. With fractional distillation, considerable help may be required, particularly in interpreting the results and deciding the probable composition of the fractions. Although the fractional crystallization and chromatographic separations are more complex to perform, students should be able to analyze their data with ease. Since it is unlikely that all of the procedures will be used with a class, several demonstrations are presented in the next activity section that may be used during the post laboratory period to show more than one separation technique. It is also recommended that other procedures be discussed at this time. For example, if students only do the filtration activity, then distillation, extraction, and/or chromatography activities may be demonstrated. Paper chromatography provides an interesting demonstration and is recommended regardless of which procedures are performed. Separations that were not done as part of the activity may be assigned as individual projects.

Extensions

Giving students the opportunity to perform one of the procedures not assigned to the entire class would be a suitable extension.

Additional laboratory activities involving separations may be found in *Laboratory Activity 1* in the *Forensic Chemistry* module. This activity uses thin layer chromatography (TLC) to separate the components in several over-the-counter drugs. This interesting laboratory activity uses a different chromatographic technique.

"Chalk Chromatography" in *Chemical Activities*, Teacher Edition (Borgford & Summerlin, 1988) is a unique chromatography activity with excellent instructions. (A student version of *Chemical Activities* is also available, but the teacher edition is sufficient if one is willing to type instructions for students.)

Assessing Laboratory Learning

1. Since all four procedures allow students to develop or practice laboratory techniques, one way to assess laboratory learning is to note which students are successful at performing the techniques involved. Hopefully, many students can successfully separate a mixture of soluble and insoluble substances by filtration.
2. On an examination, provide students with a graph showing the solubilities of several substances. Then ask students to use the data to select pairs of substances that might be separated by fractional crystallization. (To successfully use fractional crystallization, the substances to be separated should have *large* differences in solubility at *high* temperature. Secondly, the rate at which solubility changes with temperature should be quite different, *e.g.*, NaCl and KNO_3. Fractional crystallization has been used to separate NaCl, KCl, $MgCl_2$, and $MgBr_2$ from brines found in the Dead Sea between Jordan and Israel.)
3. What is one very important use of fractional distillation? (Fractional distillation is widely used in the petroleum industry to separate crude petroleum into several fractions containing useful hydrocarbon mixtures. Gasoline contains hydrocarbons that boil in a temperature range of 60-280 °C, jet fuel a range of 190-450 °C, kerosene a range of 350-550 °C, and so forth.)

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Separation of Two Immiscible Liquids

Description

This demonstration illustrates the separation of immiscible liquids. It may be used when solutions and/or other mixtures are being considered.

Materials

Separatory funnel, 60- or 100-mL
Cooking oil, 50 mL
Distilled water, 50 mL, containing 1-3 drops of food coloring

Safety

After shaking the separatory funnel it is necessary to relieve the pressure buildup by inverting the funnel and opening the stopcock. Otherwise, the stopper could "blow."

Procedure

Show students the two separate liquids. Pour them together into a large beaker. Ask students how the two could be separated completely. Pour the mixture from the beaker into a separatory funnel, stopper the funnel, and shake thoroughly. Relieve the pressure buildup. Set the funnel in a small iron ring attached to a ring stand and allow the liquids to separate. The oil layer will be the top layer. Remove stopper, then open the stopcock of the separatory funnel while holding it over the small beaker that originally contained the colored water and allow the bottom colored layer to drain out. Watch carefully so that the stopcock is closed at the precise moment when the water is completely drained (this may require some practice beforehand). Drain the remaining cooking oil into the small beaker that originally contained it. Show the class that the collected volumes are about the same as the original volumes. Figure 4 shows a sketch of a separatory funnel with the two immiscible liquids.

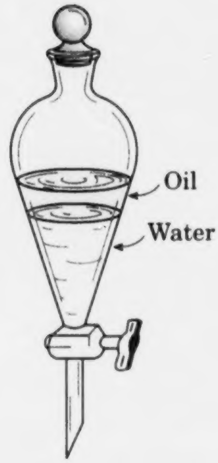


Figure 4. Separatory funnel.

Remarks

About 10 min preparation time to try the demonstration prior to performing it for students is required. 15-20 min are needed to perform the demonstration along with an explanation of what is occurring. Have colored water prepared in advance. A graduated beaker is sufficiently accurate for measuring liquid volumes. Thoroughly shake the two liquids together in the funnel while the class observes. Students may be called upon to help.

Little discussion is needed; the observations tell the story. It might be wise to indicate how the demonstration fits into the topic being considered. Ask students why the oil comprises the top layer of the two layers in the funnel *[less dense]*. How do they know it is the oil?

Demonstration 2: Fractional Distillation of Two Miscible Liquids

Description

This demonstration may be used if the fractional distillation activity is not selected for student use.

Materials

- Thermometer
- Distilling flask, 250-mL
- Boiling chips or glass beads
- Condenser
- Ringstand, utility clamps, iron ring
- Ceramic-centered wire gauze
- Hot plate
- Methanol, 50 mL
- Distilled water, 50 mL
- Safety goggles

Safety

Wear safety goggles and use a hot plate rather than an open flame. Methanol is both toxic and highly flammable.

Procedure

Mix 50 mL methanol with 50 mL distilled water. Pour the mixture into a 250-mL distilling flask and set up the distilling apparatus as illustrated in Figure 1 of *Activity 1, Part II. Fractional Distillation*. Check fittings in distillation apparatus for tightness since methanol vapors are very flammable. Heat the mixture at a temperature below the boiling point of water (100 °C/1 atm), but above the boiling point of methanol (65 °C/1 atm). The alcohol will distill first.

Remarks

See *Activity 1, Part II. Fractional Distillation* for details concerning the apparatus and procedure.

Demonstration 3: Separation of Chlorophyll Pigments by Paper Chromatography

Description

This demonstration illustrates separation of pigments by paper chromatography.

Materials

Fresh or frozen spinach (or green ivy)
Mortar and pestle
Washed sand
Solvent mixture (25 mL hexane, 25 mL ethanol, and 2 mL 6 M acetic acid)
Coffee filter paper strip about 1 inch by 8 inches (regular filter paper may be used)
Large test-tube (25- x 200-mm)
Cork to fit test-tube
Paper clip

Safety

Since the solvent mixture is highly flammable, do not have open flames nearby.

Procedure

Prepare a mixture of 25 mL hexane, 25 mL ethanol, and 2 mL 6 M acetic acid (1.7 mL conc reagent/ 5 mL solution). Grind a handful of spinach or green ivy with sand in a mortar and pestle in 25 mL of the prepared solvent. The pigments in the cells will dissolve into the solvent, giving a dark green liquid.

Pour 10 mL fresh solvent mixture into a large test-tube and place in a test-tube rack. Cut a piece of coffee filter paper about one inch wide and long enough to reach the bottom of the test-tube. Dip a pipet into the green mixture and carefully spot the pigment onto the filter paper about one inch from the end of the paper. Straighten a paper clip with one end resembling a hook. Insert the straight end into the bottom of a cork stopper and hang the strip of filter paper from the hook, the pigment dot on the end away from the hook. The pigment spot on the filter paper should be as small as possible. It may be necessary to spot several times to build up the concentration of pigment. Allow the spot to dry between applications. This might be done prior to class.

Insert the cork stopper into the test-tube so the end of the filter paper just barely touches the solvent. (*NOTE: The pigment dot must be above the level of the solvent.*) Set upright and observe the separation of the chlorophyll pigments. The same procedure may be done substituting different types of pen inks. Washable inks, particularly black, work well.

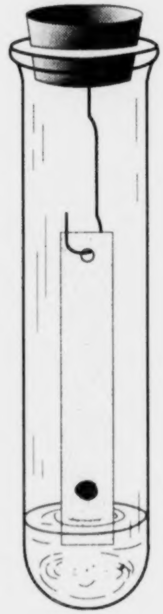


Figure 5. Paper chromatography set up.

Remarks

This demonstration may be used anywhere chromatographic separation fits in. If students do not perform a paper chromatography procedure, then the demonstration should be done.

The time needed depends upon the method used. If students help with grinding the spinach leaves and other tasks, then the demonstration will take most of a period. If you do the entire demonstration, then the entire period will be needed. In either case, the demonstration should be rehearsed prior to presenting it to a class.

NOTE: Bob Becker (Kirkwood High School, Kirkwood, MO) suggests a circular chromatography demonstration or activity using water as a solvent and Vis-a-Vis™ washable overhead projector pens. Make a pinhole in the center of a 12-cm diameter filter paper circle, fit with a cone-shaped wick of filter paper that penetrates the pinhole, and spot four colors (blue, black, red, and green are most effective) at about 1 cm from the central pinhole. Place the assembly in a Petri dish with a slightly smaller diameter than the filter paper that has about 1/2 cm of water in the bottom. Allow the water to wick up the cone-shaped central wick and chromatograph the four spots. A lovely separation pattern appears. Alternatively, you can use only a black marker in a ring. If you collect pens from several years, you will note that each has a "tag" from that year; you can ask students to see if one can distinguish between the pens from year to year.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What are the differences in properties that may be used to separate the components of a mixture? *[Differences in boiling point, particle size, solubility, densities, magnetic properties, volatility.]*
2. Name some techniques used to separate the components of a mixture. *[Simple distillation, fractional distillation, filtration, crystallization and fractional crystallization, attraction to a magnet, paper chromatography, thin layer chromatography (TLC), gas chromatography, column chromatography, high performance liquid chromatography (HPLC), gel electrophoresis.]*
3. What purpose is served in separating the components of mixtures? *[Separating the components of a mixture may serve several purposes. In order to use many of the modern instruments available (spectrometers, etc.), it is frequently necessary to have pure substances. Many analytical procedures depend upon separating mixtures to identify the components. Separation of undesired contaminants is important in the preparation of chemical products. Petroleum refining is based upon first separating the crude petroleum into different fractions (fractional distillation). The preparation of many biologically active substances requires careful separation. For example, some biologically active substances exist as enantiomers (lefthand and righthand stereoisomers), of which only one isomer is the active form. What students give as an answer will depend upon their knowledge. You will probably need to help them.]*
4. How do components of a mixture affect the properties of a mixture? *[Mixtures generally have properties that differ from those of the pure components. Some properties that are different for mixtures are melting point, boiling point, volatility, density, physiological activity, reactivity, and so forth.]*

5. How does a mixture differ from a compound? *[A mixture may be either heterogeneous or homogeneous. In either case, the mixture does not have a definite composition. For example, concentration of salt dissolved in water can have an infinite number of values. Compounds always have a definite composition as generalized by the law of definite proportions. Mixtures may be separated into their components by physical methods whereas compounds cannot.]*

6. How does a homogeneous mixture differ from a heterogeneous mixture? *[A homogeneous mixture has a uniform composition but a heterogeneous mixture does not. For example, a 1-mL sample of a 10% salt solution has the same composition regardless of where in the bulk solution a sample is obtained. Granite, on the other hand, illustrates a heterogeneous mixture. One sample of granite will have a different composition by mass from that of another sample of granite from the same source.]*

7. What are some household products that are the result of separation techniques? *[Table sugar is produced by the crystallization of a solution. Salt may be refined to remove undesired impurities, and it can be prepared by the evaporation of brine. Distilled vinegar is prepared by fractional distillation. Flavorings used in baking, candy making, beverages, perfumes, etc. are obtained by extraction. For example, fragrances used in perfumery are extracted from natural sources. Vanilla is extracted from vanilla beans using an alcohol as the extracting solvent. Decaffeinated coffee has the caffeine extracted by either a chemical process or by a "natural" process. The chemical process uses methylene chloride as the extracting solvent while the "natural" process uses water as the extracting solvent. Methylene chloride is highly toxic and a known carcinogen.]*

Extensions

A suitable activity for students is to observe labels on common household products and over the counter drugs for information about purity. (If students are sent to the store, talk to the store manager prior to sending students there to take notes.) Products to be checked include household cleansers, household ammonia, baking soda, rubbing alcohol, moth balls and moth crystals, food products such as salad oil, common analgesics such as aspirin, *etc.*

An alternate activity is to have students bring products from home or to survey the products that are present in homes. In particular, students might pay attention to such things as paint thinner, paint remover, pesticides, herbicides, *etc.* It is not recommended that pesticides, herbicides, flammable solvents, and toxic substances be brought to school. Students can copy the needed information from the labels.

Pictures in the Mind

Gas-Liquid Chromatography (GLC)

Gas-liquid chromatography is used to separate mixtures of gases and/or volatile liquids. The mixture is first vaporized and then passed through a long, narrow column packed with some finely divided solid that is coated with a nonvolatile liquid. A carrier gas, usually helium or nitrogen, sweeps the vaporized mixture through the column. The mixture vapors move along the column at different rates since they interact to different degrees with the column's liquid coating. The vapors that interact the least travel the fastest through the column while those that interact most strongly travel the slowest.

The result is that the fastest vapors exit the column first; the slowest exit last. In this way the components of the mixture are separated. At the exit end of the column, a detector detects the vapor's passing and sends along the information to a recording or memory device. When the information is displayed on a moving piece of chart paper or on a computer monitor screen, it is called a chromatogram. In this way chemists are able to "see" the number of components in a mixture.

Gas chromatography has been developed to the extent that it is possible to separate extremely complex mixtures when only nanogram (10^{-9} g) quantities are available. This development combined with advances in detectors makes the detection and analysis of trace quantities in the picogram (10^{-12} g) level possible. It is even possible to separate compounds that are different only in their isotopic composition.

A way of picturing the gas chromatography process is to imagine a large flock of seagulls (molecules) traveling down a coastline. While all are wafted along by the same tailwind (carrier gas), some stop off to feed on the ocean's (stationary phase) riches, and arrive at a given geographic point (detector) later than those gulls that did not stop or stopped only occasionally along the way. Molecules differ from seagulls in this analogy in one important respect: Chemically identical molecules ideally will behave identically with respect to the stationary phase, thus arriving at the destination (detector) in distinct clumps, or peaks.

While this analogy is simplistic, it helps visualize the chromatographic process.

Counterintuitive Examples

How pure is pure? If orange juice (pure orange juice) is filtered to remove the pulp, is it pure or a mixture? Is gasoline pure? Table salt? Iodized salt? Ivory soap—what does "99 and 44/100ths % pure" mean?

Ivory Soap is advertised to be "99 and 44/100ths percent pure." Does this mean the soap really contains 99.44 alkali metal salts of fatty acids out of every 100 "molecules"? A statement in the *Content in a Nutshell* section says that nothing is absolutely pure. So this statement about Ivory Soap means the limits of purity must be definable and measurable. In the case of Ivory Soap, when compared to three expensive imported soaps, it was found to contain fewer impurities. That is where the sales angle of "99 and 44/100ths percent pure" originated. The air mixed with the soap that lets it float is an impurity.

Most people are familiar with the "sweet" smell of air after a thunderstorm or rain. The falling water seems to remove all undesirable substances. Is it pure? Air is a mixture and not a pure substance.

Pure orange juice is another case in point. It too is a mixture although it may be made only from fresh oranges. In fact, no substance is completely pure. Impurities are always present. The impurities may be from the atmosphere, from unremoved processing materials, or even from the container holding the "pure" substance.

Modern techniques make it possible to detect impurities in substances that may have been previously considered pure. Parts per million (ppm), parts per billion (ppb), and even parts per trillion (ppt) detection has been achieved. Units of mass, molecules, or volume may be used. For example, an impurity present at 1 ppm by mass means that 1 g of material contains 1 millionth of one gram of impurity. This is 1×10^{-6} g ($1\ \mu$g) in the 1 g sample. Here is where separation techniques play an important part.

An area of active research is connected with separation techniques that allow reliable and rapid analysis of complex mixtures of pollutants found in toxic wastes, polluted streams and lakes, and contaminated biological samples. Analytical methods allowing the separation and quantitative measurement of each of the 22 individual isomers of tetrachlorodioxin at the parts per trillion range (one part in 10^{12} or 1×10^{-12} g in a 1 g sample) have been done. With such advanced techniques it is possible to pinpoint the impurities in practically any substance.

Other

Separations, instrumentation, and environmental quality —Chemistry in action.

According to *Opportunities in Chemistry Today and Tomorrow* (Pimentel & Coonrod, 1987), three chemically related questions are asked:

- "What potentially undesirable substances are present in our air, water, soil, and food?"
- "Where did these substances come from?"
- "What options are there—alternate products and processes—to reduce or remove known problems?"

Determining what substances are present in the environment requires analytical chemists to act as detectives. To do so they need to develop sensitive and selective analytical techniques. Analytical chemists deal in separations and instrumentation. Determining where the substances originate calls for the action of analytical chemists along with collaboration by meteorologists, oceanographers, geologists, climatologists, biologists, and hydrologists. The development of alternatives, however, requires that problems be tackled by the entire society.

Detection may be protection. A point that must be made is that simply detecting the presence of a substance does not mean it is a hazard. Chemists need to use their skills at detecting minute concentrations well below the hazard threshold so that frantic corrective action need not be taken. Although it seems trite, proaction rather than reaction is needed. For example, carbon monoxide is dangerously toxic at concentrations greater than 1000 ppm. This compound is always present in the atmosphere at about 1 ppm, and prolonged exposure to concentrations greater than 10 ppm is considered to produce negative health effects. Should CO be removed completely from the atmosphere? That would probably be impossible and prohibitively expensive. The real need is to be able to decide where cleanup should begin between the safe range and the known toxic threshold. This decision requires the ability to separate the CO from the other atmospheric constituents, or to develop effective instrumentation to measure it in place, thus providing protection through detection.

Some people have called for a "zero risk" approach to the environment, meaning that every single molecule of CO would be removed from the atmosphere. This goal is clearly unattainable from both scientific and economic points of view. This trend has, fortunately, been gradually replaced by a sophisticated risk assessment/risk management approach. In this approach costly crash projects are replaced by long-term fundamental environmental monitoring. Continuous monitoring requires good analytical techniques. These techniques, in turn, require improved sensitivity, selectivity, separation, sampling, accuracy, speed, and data interpretation. Separations, instrumentation, and computers are key players in such an effort.

Analytical methods to separate and measure quantitatively the 22 different isomers of tetrachlorodioxin at the parts per trillion level (1 part in 10^{12}) have been developed. The concentration of nitric acid and formaldehyde in Los Angeles smog can now be measured by infrared spectroscopy. The determination measures successfully the simultaneous concentrations of formaldehyde, nitric acid, peroxyacetyl nitrate, and ozone in the air at the parts per billion range. Although one part per billion may be too small to pose a direct hazard, its presence can be significant in atmospheric chemistry. Sulfur dioxide concentration at the ppm level in the smoke plumes of power plants has been measured. It is also possible to detect the pollutants directly from the exhaust pipe of an internal combustion engine, as in automobile exhaust.

TIPS FOR THE TEACHER

Language of Chemistry

azeotrope any mixed liquid with a boiling point lower than any liquids making up the mixture.

distillation purification process in which a liquid is evaporated and then condensed again to a liquid; used to separate dissolved solids from liquids or liquids from liquids according to boiling point.

electrophoresis movement of charged particles in a liquid or a gel caused by an electromotive force derived from a voltage applied to electrodes in contact with the liquid or gel.

filtrate liquid that passes through the filtering apparatus when a solid-liquid mixture is filtered.

filtration process used to separate an undissolved solid from a liquid.

fractional crystallization separation of two solids by using a difference in solubility at relatively low temperature; also called recrystallization.

fractional distillation separation of miscible liquids having different boiling points by heating, condensing the vapor, and collecting the distillate in fractions according to boiling point.

mixture combination of two or more substances not chemically combined and without definite proportions.

percent composition percent by mass of each component in a mixture.

polarity charge displacement in a molecule due to the different electronegativities of the component atoms, *e.g.*, polarity of water.

precipitation formation of a solid from components previously in solution.

purity degree to which a homogeneous substance consists of only one component.

residue solid that is retained in the filter paper during filtration.

solubility extent to which a solute dissolves in a solvent.

solute dissolved substance in a solution, *i.e.*, the dispersed component.

solution homogeneous mixture consisting of two or more components.

solvent dispersing or dissolving medium in a solution; generally present in greater quantity than the solute.

Pattern Recognition

Like dissolves like—polar molecules dissolve in polar solvents and nonpolar molecules dissolve in nonpolar solvents; a general rule that is probably overused since there are a number of exceptions.

Common Student Misconceptions

"'Pure aspirin' (or other product labeled 'pure') contains only aspirin."

Although common products are frequently labeled as being pure, in most cases they are not. Aspirin is made with pure aspirin, but the aspirin tablet is primarily filler (starch or other digestible substance) with 325 mg of sodium acetylsalicylate. The same is true of other analgesics such as ibuprofen (Advil®) and acetaminophen (Tylenol®).

Decision Making

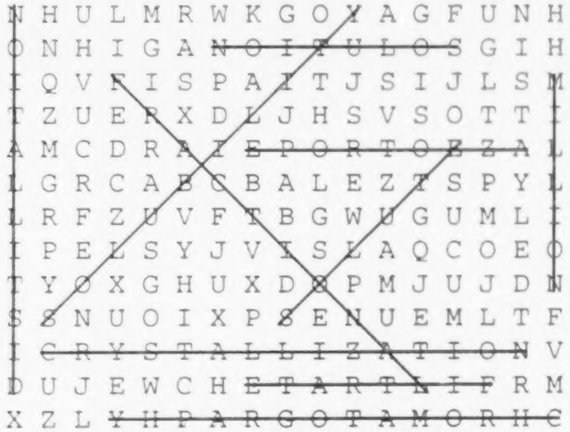
Information about decision making as it pertains to both separations and instrumentation is presented in *Decision Making* in the *Instrumentation* module.

HISTORY: ON THE HUMAN SIDE

Chromatography is a process based on differences in the rate at which the components of a mixture move through a stationary phase while under the influence of a moving phase. The word "chromatography" comes from the Greek words for "color" (khroma) and "to write" (graphein). Chromatography results in "color writing." The Russian botanist, Mikhail Tswett (1872-1920), is considered to be the first to have described how to separate mixtures of plant pigments using column chromatography. Paper chromatography was considered so important in the scientific world that its developers, Martin and Synge, received the Nobel Prize in 1952.

HUMOR: ON THE FUN SIDE

1. Word Search (see *Appendix* for master copy)



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N H U L M R W K G O Y A G F U N H
O N H I G A N O I T U L O S G I H
I Q V F I S P A I T J S I J L S M
T Z U E R X D L J H S V S O T T I
A M C D R A I E P O R T O B Z A L
L G R C A B C B A L E Z T S P Y L
L R F Z U V F T B G W U G U M L I
I P E L S Y J V I S L A Q C O E O
T Y O X G H U X D O P M J U J D N
S S N U O I X P S E N U E M L T F
I C R Y S T A L L I Z A T I O N V
D U J E W C H E T A R T L I F R M
X Z L Y H P A R G O T A M O R H C
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Type of distillation that separates miscible liquids.
2. Liquid that passes through filter paper when filtering a solid-liquid mixture.
3. The dissolved substance in a solution.
4. Separation of a solid from solution based on a difference in solubility at low temperature.
5. Process of evaporation and condensation in a separate vessel.
6. Technique used to separate colored substances in a green leaf extract.

7. A mixture of liquids with a boiling point lower than any liquid in the mixture.
8. Extent to which a solute dissolves in a solvent.
9. Homogeneous mixture of two or more components.
10. The "m" in ppm.

Answers: 1. FRACTIONAL 2. FILTRATE 3. SOLUTE 4. CRYSTALLIZATION 5. DISTILLATION 6. CHROMATOGRAPHY 7. AZEOTROPE 8. SOLUBILITY 9. SOLUTION 10. MILLION

2. See cartoons at end of module.

MEDIA

1. The *World of Chemistry* videotape "Number 22: Polymers." World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.

2. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   Viscosity of Polymer Solutions, by Gary L. Bertrand. Vol. V A, No. 2, for the Apple II computer.

3. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).

   a. For the Apple II computer running on ProDOS: AR 701

   b. For the Apple II computer: AP 702, AP 705, AP 209

   c. For IBM PCs and PC-compatibles: PC 3601, PC 4901

4. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   a. "From an Amino Acid to a Peptide Chain," "The Alpha Helix," "Hair and Silk," and "DNA Structure, Synthesis of Messenger RNA, Protein Synthesis," four chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.

   b. "Polymers," "Carbohydrates," and "Proteins," three chapters on the videodisc *Demonstrations in Organic Chemistry* (double sided, 60 min.), Special Issue 6.

5. *Elements, Compounds and Mixtures*, 20 min. video available from Coronet/NTI Film and Video, 108 Wilmot Road, Deerfield, IL 60015; (800) 621-2131; (708) 940-3640 (FAX).

EQUIPMENT AND INSTRUMENTATION

Glassware, such as adapters, condensers, funnels, flasks, beakers, *etc.* used in the procedures in this module, is available from most scientific supply houses, like Sargent-Welch, Cenco, EMD Fisher, Flinn Scientific, *etc.* This is also true of chemical substances.

Links/Connections

WITHIN CHEMISTRY

The topic relates separation of components of mixtures to different fields of chemistry.

1. Separation techniques are related to chemical and physical properties, states of matter, and phase changes. Separations are the basis for many laboratory experiences. Data analysis involves the use of mathematics. A knowledge of properties of mixtures as well as pure substances is required in designing separation techniques.

2. Industrial chemistry uses separations in quality control and monitoring of reaction streams as well as in batch and continuous processes in the manufacture of chemicals.

3. The petroleum industry is based on separation of components of crude oil by fractional distillation. A visit to a local petroleum refinery could be an excellent field trip for a chemistry class.

4. A helium refinery is another example of industrial separation of components of a mixture. Helium occurs in high percentages in natural gas deposits in the Texas panhandle near Amarillo. This is the only such refinery in the United States. Separation of helium from the natural gas involves compressing the gas mixture until the methane and other gases liquefy, leaving only helium as a gas. The helium is bled off, purified, and processed further by carefully controlling the pressure above the liquefied gas. The methane is returned to the well from which it was taken or sold outright. The concept of critical pressure and critical temperature can be applied to this technique.

5. Forensic chemistry utilizes separation techniques in determination of concentrations of various substances in the bodies of victims of crime or a person suspected of committing a crime. Illegal drugs, alcohol, and other life-threatening toxic substances are subject to investigation and require separation from other substances.

6. Medical technology laboratories in hospitals and the private sector are involved in separation techniques to determine concentrations of various substances in body tissues and body fluids. Detection of various inheritable diseases (diabetes), infectious diseases (hepatitis), and degree of injury due to a heart attack are examples of biomedical analyses that depend upon separation and instrumentation.

7. The weather bureau actually does separation techniques in determining the pollen count of the air each day.

8. Gaseous effusion was used to separate the different isotopes of naturally occurring uranium as gaseous uranium hexafluoride. By allowing the UF_6 to effuse through a series of effusion orrifices, the lighter ^{235}U isotope was successfully separated from the heavier ^{238}U isotope.

9. Electrophoresis is another recent development (after 1930). In this technique ions (electrically charged particles) move through a liquid caused by an emf (developed by a voltage source) applied to electrodes in contact with the suspension or solution. It is particularly useful in separating a mixture of proteins or the components of blood serum. It is also useful in separating DNA and RNA fragments.

10. Ultracentrifugation is another separation process that is widely used in biochemistry. In this process a mixture is placed under the influence of a centrifugal force that is many times greater than the earth's gravitational field, causing a different rate of sedimentation of different molar mass substances. Ultracentrifuges spin at rates of 5×10^4 revolutions per minute or more.

11. In order to separate the sulfur dioxide (SO_2) from the combustion products derived from burning fossil fuels, the effluent gases are passed through a piece of equipment called a scrubber. In this device, the gases are sprayed with a water suspension of lime (CaO), which converts the gaseous sulfur dioxide to solid calcium sulfite ($CaSO_3$) and calcium sulfate ($CaSO_4$).

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Since the distinction between chemistry and other disciplines, including the arts and social sciences, is rapidly disappearing, the links between chemical uses of separation and uses in other disciplines is rather broad. Medicine, in particular, makes wide use of separations in microbiology, histology, and pharmacology. (Some medical uses of separations techniques were described in *Within Chemistry*.)

TO THE CONTEMPORARY WORLD

All of the separation techniques discussed in this module are currently used by the chemical industry, in medical research, in routine medical testing, in pollution monitoring and abatement, by many agencies in the governmental and private sectors, and others. These techniques now are in widespread use and will remain so.

1. **Careers**. Analytical chemist, soil analysis, technician, medicine, environmental chemist, forensics medical examiner, medical technology, water quality control, biochemist, field biologist, petroleum engineering, teacher-professor.
2. **Community**. As mentioned previously, separation is used in many endeavors that are present and operational in most communities throughout this country as well as much of the world.
3. **Field trips**. Petroleum refinery, private analytical laboratories, water quality control board, criminology laboratories, forensics laboratory, sewer plant, metal plating industries, helium refinery, dairy industry, hospital laboratory, mining sites, pharmaceutical industries.
4. **Knowledgeable individuals**. Analytical chemists in industrial and private sectors, physicians, soil chemists, university professors, pharmacologists, forensic chemists, medical examiners, dairy chemists, environmental chemists and biologists, local water quality chemist.

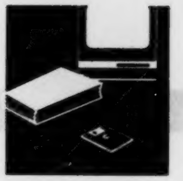
SOCIETAL (SCIENCE/ TECHNOLOGY/ SOCIETY; CURRENT EVENTS)

See *Instrumentation* module.

Extensions

Students can prepare poster sessions to be presented to the class. A poster session may include any facet of solution chemistry and separation technique employed in the private sector, industry, medicine, *etc.* The session includes poster boards illustrating the information. The presentation may last for twenty to thirty minutes. Information may be obtained from interviews, field trips, personal research, or library research.

References



Module developed by Royace Aikin and Daniel Kallus, the Texas team.

Borgford, C. L., and Summerlin, L. R. (1988). *Chemical activities.* Washington, DC: American Chemical Society.

Excellent compendium of simple, easy to perform activities. Especially recommended for this module are some of the activities in the section, *Chemical Detectives: Tools and Techniques of the Chemist.* "Chalk Chromatography" is a particularly interesting activity.

Budavari, S. (Ed.). (1989). *The Merck index* (11th Ed.). Rahway, NJ: Merck and Co.

Excellent general reference for properties, preparation, and hazards of chemicals. It also contains many useful tables in the appendix along with a formula index and cross index.

Lippincott, W. T., Gailey, K. D., Meek, D. W., and Whitten, K. W. (1984). *Experimental general chemistry.* Philadelphia, PA: Saunders.

Excellent laboratory manual that may well serve as a general reference for general laboratory procedures and instrumental techniques. The lab manual may also be used to select particular parts of experiments to use with high school classes.

McQuarrie, D. A., and Rock, P. A. (1984). *General chemistry.* New York, NY: Freeman.

This book is an excellent general chemistry reference for the teacher. It contains extensive coverage of the same topics taught in a typical high school chemistry course, but in greater depth.

Parry, R. W., Bassow, H., and Merrill, P. (1987). *Chemistry: Experimental foundations* (4th Ed.). Englewood Cliffs, NJ: Prentice-Hall.

The only CHEMS-derived text currently published. Very good information, especially how to present chemistry from an experimental point of view. The companion teacher's guide is a must.

Pimentel, G. C. (Ed.). (1963). *Chemistry: An experimental science.* San Francisco, CA: Chemical Education Material Study and W. H. Freeman.

Although it is now out of date and perhaps out of print, it still is a good reference for "seeing" how to present chemistry through "discovery." The accompanying teacher's guide is useful as are many of the experiments in the laboratory manual.

Pimentel, G. C., and Coonrod, J. A. (1987). *Opportunities in chemistry today and tomorrow.* Washington, DC: National Academy Press.

Recommended reading for all high school chemistry teachers.

Summerlin, L. R., Borgford, C. L., and Ealy, J. B. (1987). *Chemical demonstrations: A sourcebook for teachers* (Vol. 2). Washington, DC: American Chemical Society.

The demonstrations, "Colorful Mixture Separations" and "Separating Metallic Iron from Cereal" are suitable for this module (see *Food and Chemistry* module). Many of the other demonstrations are useful in any introductory chemistry course.

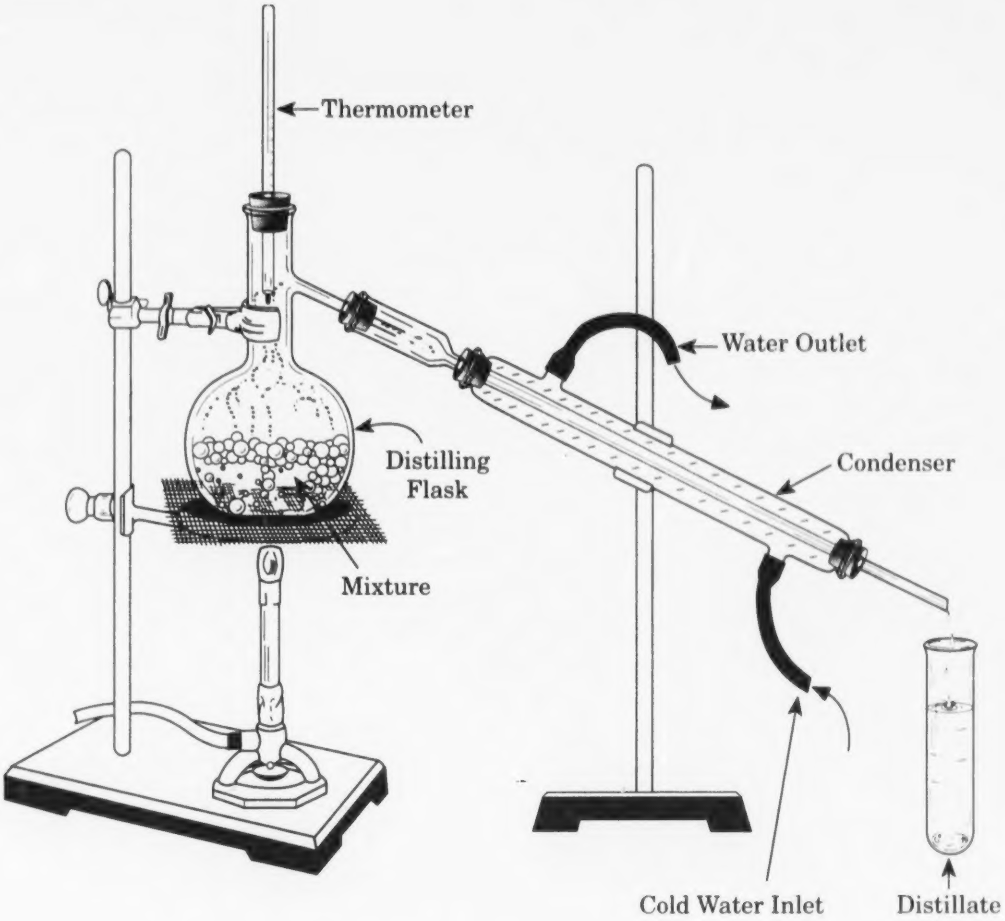
Wilbraham, A. C., Staley, D. D., Simpson, C. J., and Matta. M. S. (1987). *Chemistry laboratory manual.* Menlo Park, CA: Addison-Wesley.

A good, solid high school chemistry laboratory manual, although in a somewhat cookbook style. It presents useful experiments including a good presentation of laboratory safety and laboratory techniques.

Appendix

- **Transparency Masters**
  1. A Simple Distillation Apparatus
  2. Typical Apparatus Used in Separation Techniques
  3. Word Search
- **Humor**

A Simple Distillation Apparatus

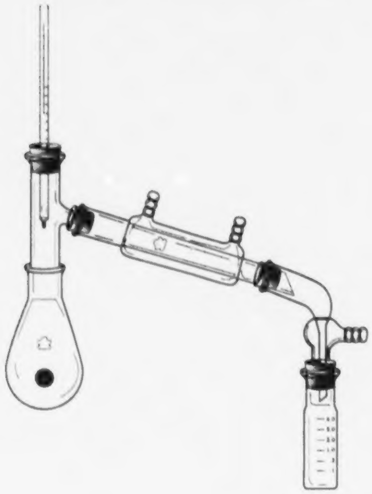

Thermometer
Water Outlet
Distilling Flask
Condenser
Mixture
Cold Water Inlet
Distillate

80

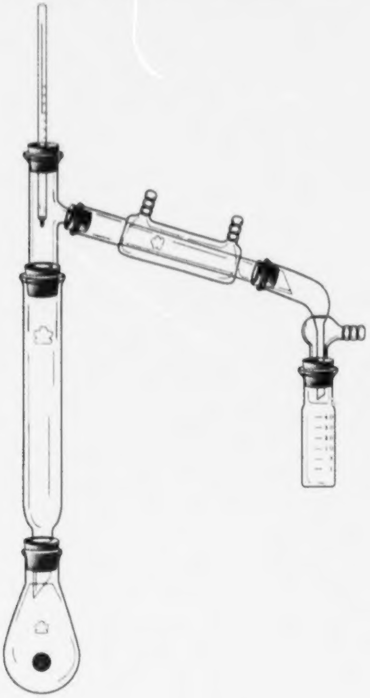
Typical Apparatus Used in Separation Techniques

(Courtesy of Kontes Glass Company)

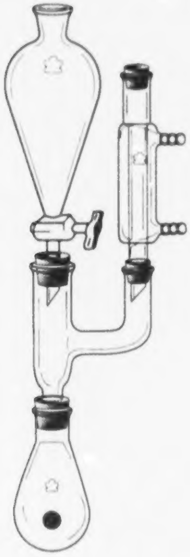
Simple Distillation



Fractional Distillation

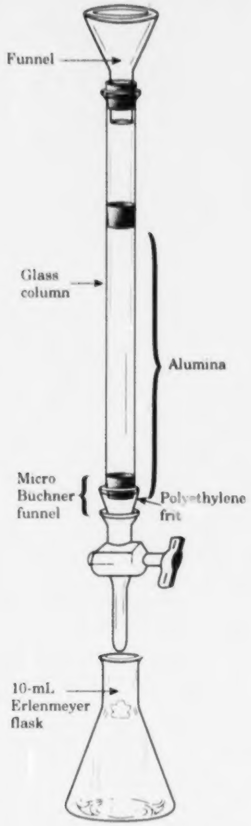


Reflux with Addition

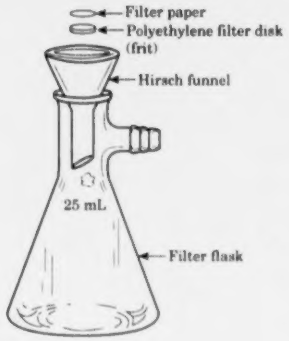


Typical Apparatus Used in Separation Techniques Cont.

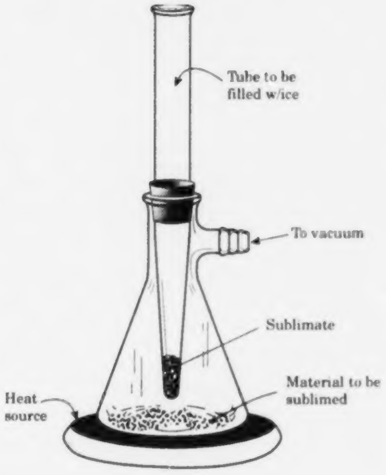
Column Chromatography


Funnel
Glass column
Alumina
Micro Buchner funnel
Polyethylene frit
10-mL Erlenmeyer flask

Hirsch Funnel Filtration


Filter paper
Polyethylene filter disk (frit)
Hirsch funnel
25 mL
Filter flask

Sublimation

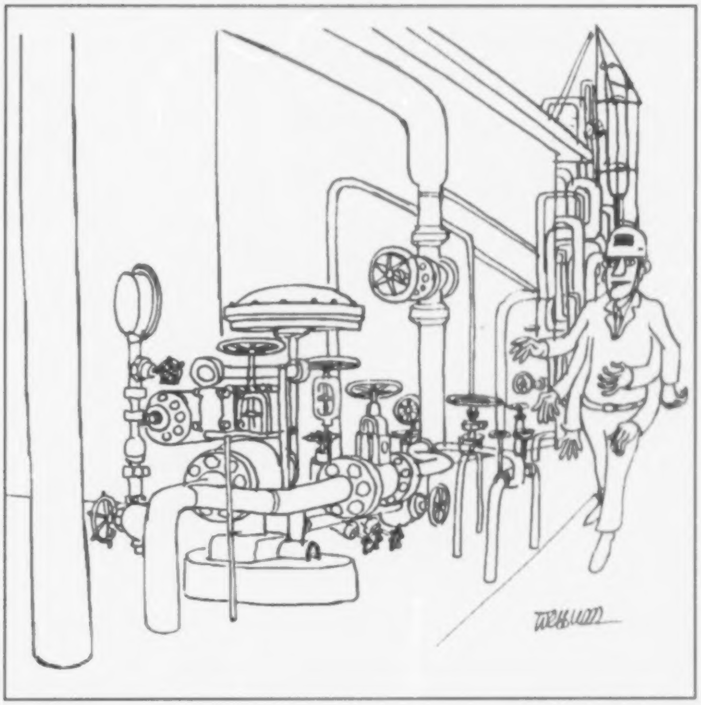

Tube to be filled w/ice
To vacuum
Sublimate
Material to be sublimed
Heat source

Word Search

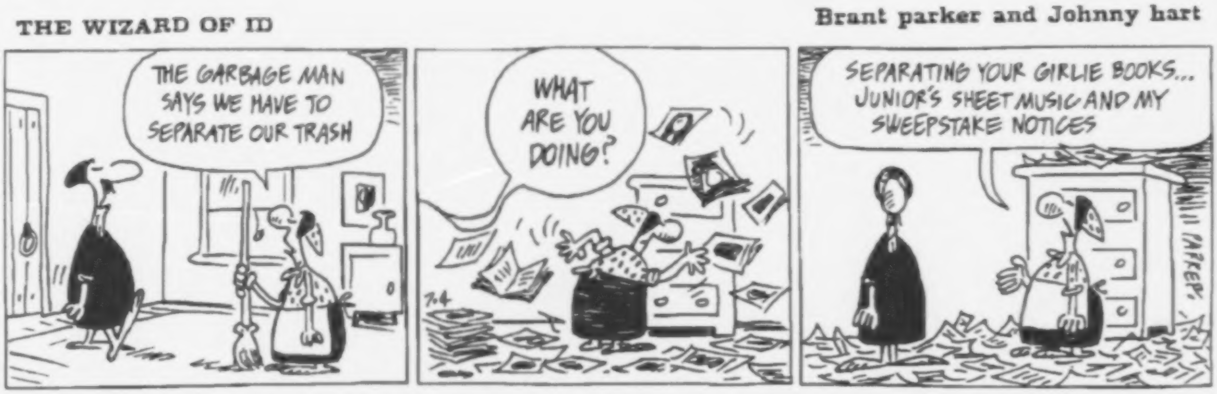
N H U L M R W K G O Y A G F U N H
O N H I G A N O I T U L O S G I H
I Q V F I S P A I T J S I J L S M
T Z U E R X D L J H S V S O T T I
A M C D R A I E P O R T O E Z A L
L G R C A B C B A L E Z T S P Y L
L R F Z U V F T B G W U G U M L I
I P E L S Y J V I S L A Q C O E O
T Y O X G H U X D O P M J U J D N
S S N U O I X P S E N U E M L T F
I C R Y S T A L L I Z A T I O N V
D U J E W C H E T A R T L I F R M
X Z L Y H P A R G O T A M O R H C

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Type of distillation that separates miscible liquids.
2. Liquid that passes through filter paper when filtering a solid-liquid mixture.
3. The dissolved substance in a solution.
4. Separation of a solid from solution based on a difference in solubility at low temperature.
5. Process of evaporation and condensation in a separate vessel.
6. Technique used to separate colored substances in a green leaf extract.
7. A mixture of liquids with a boiling point lower than any liquid in the mixture.
8. Extent to which a solute dissolves in a solvent.
9. Homogeneous mixture of two or more components.
10. The "m" in ppm.



CHEMTECH, March 1972, p. 190. © American Chemical Society. Reprinted with permission.


THE WIZARD OF ID
Brant parker and Johnny hart
THE GARBAGE MAN SAYS WE HAVE TO SEPARATE OUR TRASH
WHAT ARE YOU DOING?
7-4
SEPARATING YOUR GIRLIE BOOKS... JUNIOR'S SHEET MUSIC AND MY SWEEPSTAKE NOTICES
PARKER

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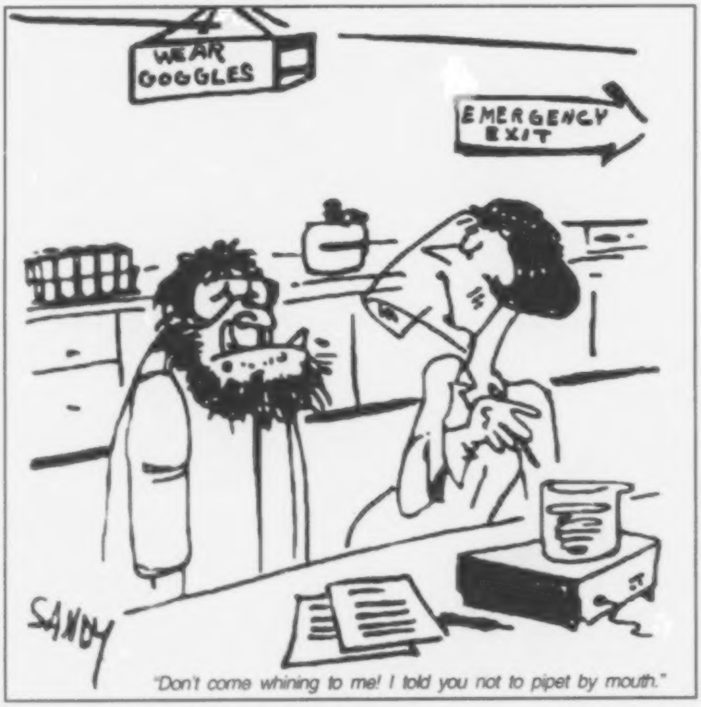
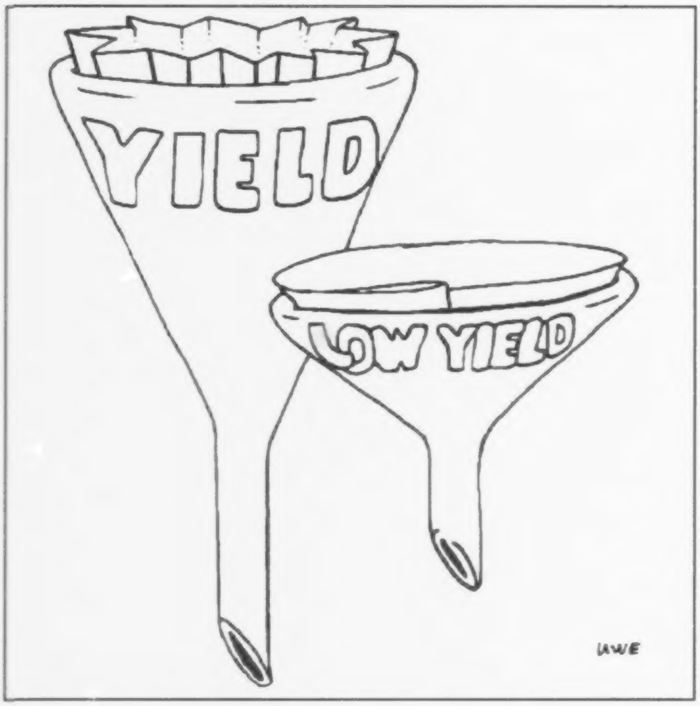

WEAR GOGGLES
EMERGENCY EXIT
SANDY

"Don't come whining to me! I told you not to pipet by mouth."

CHEM 13 NEWS, November 1982, p. 7. Reprinted with permission.


YIELD
LOW YIELD
UWE

CHEM 13 NEWS, April 1982, p. 5. Reprinted with permission.

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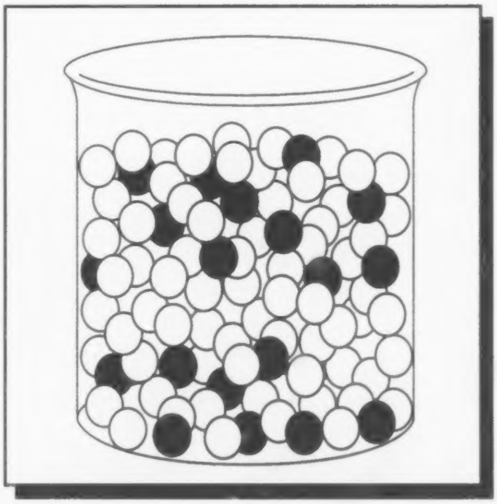


SOLUTIONS

ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Solutions make up a large, important part of chemistry. Usually solutions are considered in terms of the ability of liquids to dissolve solids, other liquids, or gases, forming a homogeneous mixture. However, solutions can be solids in solids, liquids in solids, gases in solids, solids in liquids, liquids in liquids, gases in liquids, and gases in gases. The major component of a solution is termed the solvent and the other component(s) is(are) termed the solute(s). In some cases these terms are arbitrary.

A somewhat arbitrary distinction is often made between homogeneous solutions and colloids, which are solid, liquid, or gaseous particles (made up of particles large enough to diffract or scatter light—the Tyndall effect) dispersed in solid, liquid, or gaseous media. The dispersed particles are called colloids. The thermal motion (Brownian movement) of such small particles is sufficient to keep them from settling out in the earth's gravitational field.

"Like dissolves like" is an expression that qualitatively expresses the experimental observations that polar and ionic substances dissolve in polar solvents and nonpolar substances dissolve in nonpolar solvents—*other things being equal*. A solvation model that considers interactions between polar solvent molecules and polar and ionic solutes is helpful in considering the solvation properties of water and other polar solvents. However, some aqueous solutions, and most solutions involving nonpolar solvents, depend on the randomness (entropy increase) obtained in the solution process.

Solution concentrations (*e.g.*, mol/L) should be distinguished from amounts (*e.g.*, moles). Several methods for expressing solution concentrations are important—reaction stoichiometry in solution is normally expressed in terms of the molarity (M) of the reacting species, whereas colligative properties are compared using either molality (m) or mole fractions (X), and impurities are often quoted in parts per million (ppm) or parts per billion (ppb). Relative concentrations are often expressed as dilute or concentrated, or by the terms unsaturated, saturated or supersaturated.

Solution components can be detected and evaluated by a variety of methods and can often be separated by distillation, crystallization, or chromatography. Colligative properties are useful in determining the nature of solutes in solutions. These properties together with conductivity have allowed scientists to quantify electrolytes (ionic solutes) in solutions. Solutes in water are often categorized as strong electrolytes, weak electrolytes, and non-electrolytes, reflecting the conductivity of the substance when dissolved.

PLACE IN THE CURRICULUM

This module should follow a discussion of the states of matter, changes of state, moles, and stoichiometry. It provides a good vehicle for either introducing or augmenting a discussion of solution stoichiometry.

CENTRAL CONCEPTS

1. The designations solute and solvent can be usefully made in many types of solutions (*Demonstration 1*).
2. Various types of solutions are possible, based on whether the intermixed species are solids, liquids, or gases (*Demonstration 1*).

3. A solution is a homogeneous mixture of two or more substances where the solvent and solute particles are of typical molecular dimensions (*i.e.*, < 1 nm). A mixture made up of larger particles is considered a colloidal suspension, which is observable via the Tyndall effect (*Demonstration 2*).
4. Like dissolves like, in general. Nonpolar solvents tend to dissolve nonpolar substances and polar solvents tend to dissolve polar and ionic substances.
5. A microscale model of the solution process within a liquid solvent can be constructed based on the principle of solvation.
6. Concentration (*e.g.*, mol/L) should be distinguished from amount (*i.e.*, moles).
7. Common quantitative methods for expressing solution concentration include percent by mass, percent by volume, molality (m), molarity (M), mole fraction (X), and parts per million (ppm).
8. Stoichiometry in solutions is conveniently based on molar concentrations (M) of the reacting species (*Laboratory Activity 1*).
9. Relative concentrations of a solute/solvent system can often be expressed by the terms dilute and concentrated, or by the terms unsaturated, saturated, and supersaturated.
10. A solution can be separated into its components through processes such as distillation, crystallization, and chromatography (*Demonstration 3*).
11. Colligative properties of solutions include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure (*Laboratory Activity 2*).
12. Solutes in water are often categorized as either electrolytes (strong, if completely ionized in water or weak, if only partially ionized) or non-electrolytes (nonionized).
13. Generally, gases become *less* soluble in liquid solvents as the temperature is increased. On the other hand, most (but not all) solids and liquids become *more* soluble in liquid solvents as the temperature is increased.

RELATED CONCEPTS

1. The Mole
2. Gases
3. Condensed States
4. Stoichiometry

RELATED SKILLS

1. The ability to measure mass, volume, and temperature.
2. The ability to write and balance chemical equations.
3. The ability to interpret graphs.

PERFORMANCE OBJECTIVES

After completing their study of solutions, students should be able to:

1. distinguish between solutes, solvents, and solutions.
2. list the multitude of solution types that exist.
3. discuss the difference between solutions and colloids.

4. apply the generality that polar solvents tend to dissolve polar substances and ions, whereas nonpolar solvents tend to dissolve nonpolar species.
5. explain the distinction between concentrations and amounts.
6. calculate the concentration of solutions using molarity and molality. (Calculations using the other units are optional.)
7. define: dilute, concentrated, unsaturated, saturated, and supersaturated.
8. explain how distillation, crystallization and chromatography are used to separate solutions into their components.
9. discuss colligative properties including vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.
10. describe the differences between electrolytes (both weak and strong) and non-electrolytes.
11. explain the temperature dependence of gas solubility in liquid solutions.
12. explain the temperature dependence of liquid and solid solubility in liquid solutions.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Limiting Reactant for a Chemical Reaction in Solution

Introduction

Many chemical reactions are conducted by combining solutions of reagents that will react rather than by reacting the pure materials. This strategy is particularly useful when the reactants are pure solids. Also, when synthesizing a compound it is common to have an excess of one of the reactants present in order to assure complete reaction of the second reactant. In such a reaction, the reactant present in excess is called the excess reactant, and the reactant that limits the extent of the reaction is called the limiting reactant (see *Stoichiometry* module). It is important to know which solution placed into the reaction vessel is providing the limiting reactant in order to know the maximum amount of product (theoretical yield) that can be obtained.

Purpose

To determine the limiting reactant in a chemical reaction and the stoichiometric ratio in the balanced equation.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. If you should come into contact with any reagents you should immediately wash the affected area with large amounts of water.
3. Be sure to operate the centrifuge according to the procedure given by your teacher. Be sure to place a second test-tube containing the same amount of material in a test-tube holder *opposite* the holder containing the sample. Do not attempt to stop the centrifuge by making contact with the spinning head. If a test-tube should break while in the centrifuge, notify your teacher at once—do not attempt to clean the centrifuge without direction from your teacher and only after the centrifuge has been disconnected from the electrical outlet.
4. Be careful not to contact the hot surfaces of the oven when placing objects into or removing objects from the oven. Also, objects being removed from the oven are very hot and should be handled with tongs or other appropriate utensils as demonstrated by your teacher.
5. Do not pour any substances from this experiment down the drain. Dispose of them as indicated by your teacher.

Procedure

1. Obtain four 13 x 100-mm test-tubes or other tubes usable as centrifuge tubes, two 50-mL beakers, two 250-mL beakers, one stirring rod, three Beral™ pipets, and a wash bottle filled with distilled water.
2. Place 10 mL 1.0 M $BaCl_2$ solution into a 50-mL beaker and label the beaker. Place 15 mL 0.50 M Na_2SO_4 solution into another 50-mL beaker and label the beaker.

3. Take clean and dry test-tubes and label the test-tubes 1, 2, 3, and 4 using a pencil. Measure and record the mass of each of the test-tubes to the nearest 0.01 g.

4. Use one of the Beral™ pipets to measure 2.0 mL 1.0 M $BaCl_2$ solution from your beaker into each of the four numbered test-tubes. *Be careful when filling the pipet to avoid air bubbles in the liquid column in the pipet. Use one of the 250-mL beakers as a test-tube rack. Make sure that the pipet used with the* $BaCl_2$ *solution stays in the beaker containing the* $BaCl_2$ *solution after you have finished using it.*

5. Use a second Beral™ pipet to measure volumes of 0.50 M Na_2SO_4 solution indicated below into the test-tubes. Expel the Na_2SO_4 solution into the test-tubes in a fast stream to promote mixing of the solutions and make sure that the tip of the pipet does not go into the solutions in the test-tubes.

6. Place 2.0 mL 0.50 M Na_2SO_4 into test-tube number 1, 3.0 mL into test-tube 2, 4.0 mL into test-tube 3, and 5.0 mL into test-tube 4. *Make sure that the Beral™ pipet used for the* Na_2SO_4 *solution stays in the beaker containing the* Na_2SO_4 *solution after you have finished using it.*

7. Stir the solutions with the stirring rod; rinse the stirring rod with a stream of distilled water from your wash bottle each time it is withdrawn from a test-tube making sure that the washings go into the test-tube from which you are withdrawing the rod.

8. Using water from your wash bottle fill all the test-tubes to the same height. The liquid should be about 2 cm from the top of the test-tubes. Do not overfill! Place the test-tubes into the centrifuge using two of them to counterbalance the other two and spin them for about one minute. Remove the test-tubes from the centrifuge. *The solid* $BaSO_4$ *should be at the bottom of the test-tube, and the supernatant liquid (the liquid above the precipitated solid) should be clear, or slightly cloudy. If the liquid is very cloudy, check with your teacher to determine if you should continue to centrifuge the tubes. Your teacher may also recommend additional steps to aid the settling of the precipitate.*

9. Use the third Beral™ pipet to remove the supernatant liquid from above the precipitates and discard the liquid into the 250-mL beaker not being used as a rack. *Be careful not to remove any of the settled precipitate from the test-tubes.*

10. Fill each test-tube about 2/3 full with distilled water from your wash bottle. Stir each solution with the stirring rod so that the settled precipitate is dispersed into the liquid. Use the wash bottle to rinse the stirring rod as you did in Step 7. Now add distilled water from your wash bottle to the test-tubes to make the level of liquid the same in each test-tube, but at least 2 cm from the top. Centrifuge the tubes as you did in Step 8. *Once again, the solid should be settled and the supernatant liquid clear.* Remove the supernatant liquid from the test-tubes with the third Beral™ pipet taking care not to remove any of the solids.

11. Place the test-tubes containing the wet solids into a 250-mL beaker and place the beaker into the drying oven. *Be sure it has an ID mark, such as your initials, on it.*

12. When the solids have dried remove them from the oven. *CAUTION: The beaker is* ***hot*** *and should be handled with tongs or other approved handling equipment.* Place the beaker on a wire gauze on your desk and allow the beaker, the test-tubes and the solids to cool. Measure and record the mass of each of the test-tubes to the nearest 0.01 g.

13. Dispose of the test-tube contents, the excess reagents, and waste solutions as directed by your teacher. (Do not pour the chemicals down the drain!) Wash all glassware and the pipets and return all items to the appropriate place.
14. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

1. Make a table with the following column headings (use the long dimension of your paper for the top of the table): [1] Test-tube number, [2] Mass tube (g), [3] mL 1.0 M $BaCl_2$, [4] mL 0.50 M Na_2SO_4, [5] Mass tube + $BaSO_4$ (g), [6] Mass $BaSO_4$ obtained (g), [7] Moles $BaSO_4$ obtained, [8] Moles $BaCl_2$ added, [9] Moles Na_2SO_4 added, [10] Moles $BaSO_4$ expected if all $BaCl_2$ reacted, [11] Moles $BaSO_4$ expected if all Na_2SO_4 reacted, and [12] Limiting reactant.
2. Fill in columns [1] through [5] from information from your data sheet.
3. Calculate the mass and moles $BaSO_4$ obtained in each test-tube and record the values in columns [6] and [7], respectively.
4. Calculate the moles $BaCl_2$ and Na_2SO_4 added to each test-tube and record the values in columns [8] and [9], respectively.
5. Write a balanced equation for the reaction that occurred and use information from columns [8] and [9] to calculate the moles $BaSO_4$ expected if (a) all the $BaCl_2$ added had reacted and (b) all the Na_2SO_4 reacted for each of the test-tubes, and record the values in columns [10] and [11], respectively.
6. Compare the values in columns [7], [10] and [11] and decide which of the reactant solutions, $BaCl_2$ or Na_2SO_4, contained the limiting reactant and indicate your choice in column [12].

Implications and Applications

1. Does the solution containing the lower concentration of reactant always contain the limiting reactant? How do the experimental results verify your conclusion?
2. Does the smaller volume of solution always contain the limiting reactant? How do the experimental results verify your conclusion?
3. What pieces of information do you need before calculating the maximum amount of product to be expected from a reaction where the reactants are in solution?
4. Corn sweetener is a sweetening agent obtained from corn. A recipe calls for 8.0 ounces of pure corn sweetener. You go to the supermarket and find that there are two brands of corn sweetener both in solution form. Brand A comes in a 16 ounce bottle, is 40% corn sweetener by mass and costs \$ 0.59. Brand B comes in a 24-ounce bottle, is 45% corn sweetener by mass, and costs \$ 0.95. Which brand costs less per ounce of pure corn sweetener? How many ounces of the less expensive brand should be added to the recipe in order to get the 8.0 ounces of pure corn sweetener?
5. Barium ion in solution is highly toxic if ingested. Why can large quantities of $BaSO_4$ be ingested by individuals undergoing gastrointestinal X-rays without any apparent ill effects?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Limiting Reagent for a Chemical Reaction in Solution

Major Chemical Concept

The major chemical concept is the stoichiometry involving solutions and the determination of the limiting reactant in a chemical reaction.

Level

Honor students.

Expected Student Background

Students should be capable of accurately weighing items on whatever balance is being used. A balance such as a Centigram™ balance should suffice for this experiment. The correct method of using a centrifuge should be demonstrated (or reviewed if used before). Related concepts include balancing equations, molarity, and stoichiometric calculations and the use of limiting reactant in such calculations.

Time

Because of a bottleneck that can occur with using the centrifuge, and the need to allow about 20 min for drying precipitates in the drying oven and 5-10 min needed for cooling, about 80 min is required to complete this experiment. The experiment can be conveniently broken into two parts with Steps 1 through 10 being conducted one day and the experiment being completed the second day. If you intend to complete the experiment in one day you should plan to conduct the pre-laboratory discussion the day before the experiment. Teacher/technician preparation time for the experiment is about 45 min, mainly devoted to making the two required solutions and calibrating the Beral™ pipets.

Safety

Read the *Safety Considerations* in the *Student Version*. Barium ion in solution, or in a soluble compound, is *highly toxic if ingested;* therefore, students should wash their hands thoroughly before leaving the laboratory. Barium compounds must *not* be discarded down the drain. There should be a waste container for these salts and solutions. After all students have completed the activity, you can convert the sample to *very* insoluble barium sulfate by adding excess sodium sulfate. The barium sulfate waste can be filtered, dried, and stored for proper disposal.

Safety precautions related to using a centrifuge should be stressed—be sure to place a second tube containing the same amount of material in a tube holder *opposite* the holder containing the sample. Because students may be in a hurry at the time when they are using the centrifuge, stress that the centrifuge should not be manually stopped by the insertion of fingers or other objects. If safety considerations don't impress students you may point out that quickly stopping a centrifuge tends to bring the solid back up into the liquid, thereby undoing what centrifuging has just accomplished. Also warn students to keep ties and long hair away from the spinning centrifuge. Occasionally a test-tube breaks while being centrifuged. This usually does not result in flying glass but students should be given instructions about cleanup procedures.

Probably the most common accident involves finger burns when placing glassware into or removing it from the drying oven or handling glassware that has not sufficiently cooled. Proper techniques for handling hot glassware should be reviewed before the experiment.

Materials (For 24 students working in pairs)

Balances, Centigram™ or equivalent
Centrifuge
Drying oven

Nonconsumables

48 Test-tubes, 13 x 100-mm or larger to fit available centrifuge
24 Beakers, 50-mL
24 Beakers, 250-mL
36 Plastic Beral™ pipets, stems calibrated to 1.0 mL
12 Stirring rods
12 Wash bottles filled with distilled water

Consumables

1.0 M Barium chloride, $BaCl_2 \cdot 2H_2O$, 200 mL (48.8 g $BaCl_2 \cdot 2H_2O$ per 200 mL)
0.50 M Sodium sulfate, $Na_2SO_4 \cdot 10H_2O$, 250 mL (40.3 g $Na_2SO_4 \cdot 10H_2O$ per 250 mL)

Advance Preparation

Solutions should be prepared as above. If the Na_2SO_4 in your stockroom is in the anhydrous form, 17.8 g per 250 mL yields the desired 0.50 M solution.

All glassware should be clean and dry. If students have to clean and dry the glassware, particularly the test-tubes, before beginning they will not have sufficient time to complete the activity. If the plastic Beral™ pipets are to be reused they should also be carefully washed to avoid cross-contamination during reuse and all visible traces of moisture removed.

The Beral™ pipet stems may be calibrated for 1 mL by filling a 10-mL graduated cylinder to the 10 mL mark with distilled water and drawing sufficient water into the pipet to bring the water in the cylinder to the 9 mL mark. The water level in the pipet stem should then be marked with a waterproof felt marker.

The drying oven should be turned on at the beginning of the class so that it comes to drying temperature (120-130 °C) by the time it is to be used. The thermostat settings should be determined the day before the laboratory.

Pre-Laboratory Discussion

The pre-laboratory discussion should take place the day before the activity is to be performed if it is to be completed in one 80-min period.

1. Review safety considerations and disposal methods involved with the handling of the chemical substances being used.
2. Emphasize the necessity not to mix the Beral™ pipets during the activity—one should be used for the barium chloride solution and another used for the sodium sulfate solution, and the third for removal of the supernatant liquid.
3. Make sure that students understand the difference between *amount* and *concentration* of a solution.
4. Review mole and molarity calculations as well as reaction stoichiometry calculations. The net-ionic equation for the reaction in this experiment is:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

The overall reaction is:

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + Na_2SO_4(aq)$$

Therefore, one mole of $BaCl_2$ reacts with one mole of Na_2SO_4. You should be able to elicit this information from your students if they have studied chemical stoichiometry.

5. Review the procedure used and safety considerations in operating the centrifuge and the drying oven.
6. Explain that students can use the calibration marks on the beakers to measure the 10- and 15-mL volumes needed in Step 2 as only approximate volumes are necessary.
7. Because the test-tubes are to be put in the oven, labeling in Step 3 should *not* be made with a wax marking crayon or with paper labels.

Teacher-Student Interaction

One problem that arises is the fineness of the barium sulfate precipitate. Even after extended centrifuging the precipitate may not settle; some precipitate may adhere to the sides of the test-tube. The adherence to the test-tube is not a problem, since the precipitate will remain in the test-tube and be weighed. The precipitate that does not settle is a very small fraction of the total precipitate and, if removed with the supernatant liquid, will cause a small percent error. However, you may want to tell students to add a small quantity of *soap* (not detergent) solution by dipping the stirrer in a soap solution, mixing the solution with the stirrer, and then recentrifuging. This will usually aid the settling of the precipitate.

Monitor student use of the centrifuge. Any large vibration usually signals that the centrifuge is not balanced. When in a hurry, students tend to disregard the admonition against using fingers or other objects to stop the centrifuge.

Monitor the use of the drying oven. Students may need help in removing the hot beaker and test-tubes. It is best if all students place their beakers into the oven within minutes of each other. This strategy permits the oven to heat up to the required drying temperature and remain there, speeding drying.

Circulate around the room while students are doing their weighings. It is important that the weighings be done properly and masses recorded to 0.01 g.

Anticipated Student Results

1	2	3	4	5	6	7	8	9	10	11	12
T.T. No.	T.T. Mass (g)	Reactant Vol. 1.0 M $BaCl_2$ (mL)	0.50 M Na_2SO_4 (mL)	Mass of T.T. + $BaSO_4$ (g)	Mass of $BaSO_4$ Obtained (g) (5-2)	Moles of $BaSO_4$ Obtained $x10^{4*}$	Moles $BaCl_2$ Added $x10^4$	Moles Na_2SO_4 Added $x10^4$	Moles $BaSO_4$ to Expect if all $BaCl_2$ Reacted $x10^4$	Na_2SO_4	Limiting Reactant
1	7.75	2.0	2.0	7.95	0.20	8.6	20	10	20	10	Na_2SO_4
2	8.66	2.0	3.0	8.97	0.31	13	20	15	20	15	Na_2SO_4
3	7.55	2.0	4.0	7.99	0.44	19	20	20	20	20	Either
4	7.50	2.0	5.0	7.95	0.45	19	20	25	20	25	$BaCl_2$

* The actual yield should be close to the theoretical yield. The factor 10^4 means that the quantities in columns 7-11 have been multiplied by that amount to yield numbers in the tens range. The actual quantity is obtained by reversing the process and dividing by 10^4; *e.g.*, 8.6 = 8.6×10^{-4} (in Column 7; see *Introductory Module* for use of exponents).

Answers to Implications and Applications

1. No. Sodium sulfate was the solution of lower concentration, but the results in Tube 4 show that barium chloride was the limiting reactant.
2. No. There were 2.0 mL barium chloride solution and 3.0 mL sodium sulfate solution in Tube 2 and sodium sulfate was the limiting reactant.
3. You must know the stoichiometric ratio from the balanced equation, the concentrations of the reactants and the volumes of solutions mixed together in order to determine the limiting reactant and consequently the maximum amount of product to be expected.
4. Brand A costs \$ 0.092 per ounce of corn sweetener. Brand B costs \$ 0.088 per ounce of corn sweetener. Brand B (17.8 ounces) has to be used to obtain 8.0 ounces of corn sweetener.
5. Barium sulfate is so insoluble that a negligible amount of barium ions go into solution. The barium sulfate ingested in "liquid" during a gastrointestinal X-ray is really a slurry of barium sulfate in water similar to the slurry obtained in this activity when the barium sulfate precipitate is washed with distilled water.

Possible Extensions

Depending upon the students' ability, the topics covered prior to the *Solutions* module and the time allotted, the following topics might be discussed:

1. The overwhelming number of chemical processes that take place in solution.
2. Introduction to the use of stoichiometric ratios in quantitative analysis.
3. The need to introduce excessive concentrations of reactants to shift equilibrium.

Assessing Laboratory Learning

1. Laboratory Practical. The use of the centrifuge may be tested as part of a laboratory practical. Also, an "unknown" concentration of one of the solutions in this experiment may be presented and students asked how they would go about determining the concentration of that solution—or asked to actually do it.
2. Written Examination. Problems involving the use of solutions and limiting reactant determination could be presented. Students could be asked to construct a "pictures in the mind" illustration of the limiting reactant for the $BaCl_2$-Na_2SO_4 reaction similar to the H_2-O_2 reaction illustrated in the *Stoichiometry* module.

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: The Freezing Point Depression of Lauric Acid

Introduction

Dissolved solute particles disrupt the crystallization process and lower the freezing point of the solvent. The amount the freezing point is depressed is directly proportional to the number of moles of particles dissolved.

Purpose

To determine the freezing point of lauric acid and to observe the effect a solute has on the freezing point.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Proper laboratory safety procedures must be followed.
3. The burners heat metal and glass, so proper care must be taken not to burn yourself. Hot and cool glass look alike.
4. Any substances spilled on you should be immediately washed off with large amounts of water and the incident reported to your teacher.

Procedure

Part I: The Freezing Point of Lauric Acid

1. Prepare one hot water bath (65 °C) and one warm water bath (30 °C). Set up two ring stands and place 400 mL water into each 600-mL beaker; using the burner, heat one bath to 65 °C and warm the other bath to 30 °C.
2. Weigh the test-tube (25 x 200-mm) empty; fill the test-tube to a height of about 5 cm with solid lauric acid and weigh the test-tube and lauric acid.
3. Clamp the test-tube with the lauric acid in the hot water bath and allow the acid to melt.
4. When the lauric acid has just melted, remove the test-tube from the water bath and place the thermometer and stirrer into the lauric acid.
5. Place the test-tube containing the melted lauric acid, thermometer and stirrer into the 30 °C water bath and clamp the test-tube in place. Stir the lauric acid and record the temperature every 30 seconds until the temperature of the lauric acid has dropped to 40 °C. *CAUTION: Do not attempt to remove or even move the thermometer or stirrer from the solidified lauric acid!*

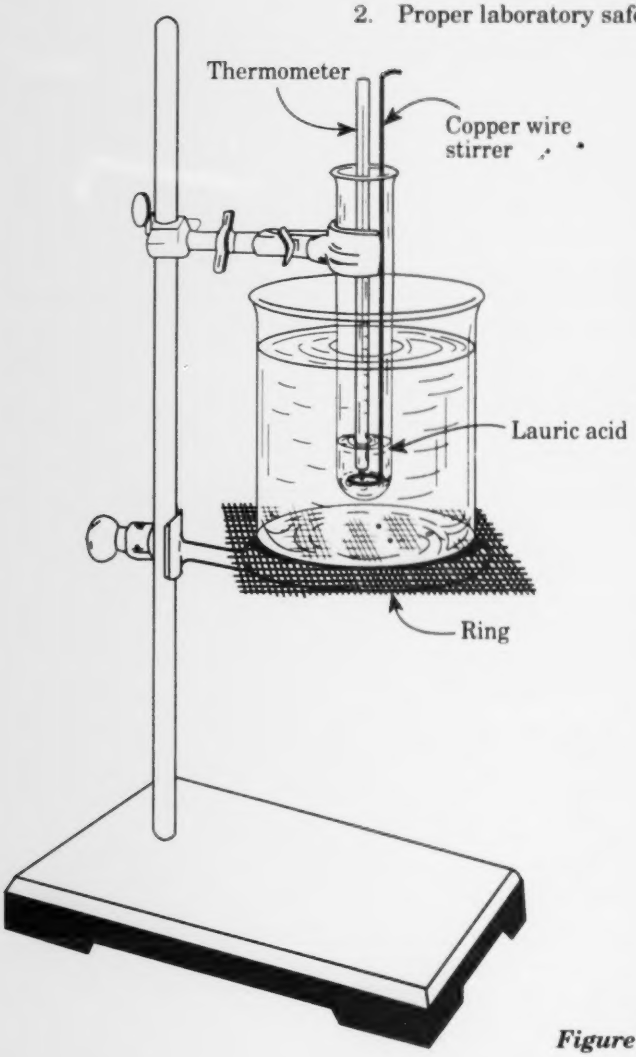


Figure 1. Elevated-temperature freezing-point apparatus.

Part II: The Freezing Point Depression of Lauric Acid

6. Weigh 2.5 g benzoic acid.
7. Remelt the lauric acid as you did in Step 3 above, and as soon as the lauric acid is melted, dissolve the benzoic acid in the lauric acid. Take care when adding the benzoic acid to make sure that all of it adds to the lauric acid, and that none of it sticks to the sides of the test-tube. You will need to stir the mixture vigorously with your stirrer in order to get the benzoic acid to dissolve completely.
8. Place the test-tube containing the melted lauric acid, benzoic acid, thermometer and stirrer into the 30 °C water bath and clamp the test-tube in place. Stir the mixture and record the temperature every 30 seconds until the temperature of the mixture has dropped to 36 °C.
9. Melt the mixture in the hot water bath and carefully remove the thermometer and stirrer from the mixture. Quickly wipe the thermometer and stirrer with a few layers of paper towels before the mixture has had a chance to solidify onto the thermometer.
10. Dispose of the paper towels and the test-tube contents as directed by your teacher, and return all materials to their proper places.
11. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

1. Graph the temperature *vs.* time for each trial.
2. What is the freezing point of lauric acid? Explain your reasoning.
3. What is the freezing point of the mixture of lauric acid and benzoic acid? Explain your reasoning.
4. How many moles of benzoic acid were used?
5. Calculate the molality (mol of benzoic acid/kg of lauric acid) of the solution.
6. Calculate the change in the freezing point temperature in going from the pure solvent to the solution.
7. Calculate the freezing point depression constant (K_f) and compare this with class results. K_f = change in freezing point divided by molality of solution. Remember that molality is moles of solute per 1000 g of solvent.

Implications and Applications

1. Is the freezing point the same as the melting point for pure lauric acid? Why?
2. Explain what is happening to the energy removed from the lauric acid in each part of the graph for the cooling of the pure lauric acid.
3. What is happening, on the molecular level, during the time when the graph of the pure lauric acid is a flat, straight line?
4. Why is the graph of the pure lauric acid different from the graph of the mixture of acids?
5. What is happening, on the molecular level, during the time when the graph of the lauric acid plus benzoic acid shows a minimal temperature change with time?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: The Freezing Point Depression of Lauric Acid

Major Chemical Concept

The major chemical concept is the freezing point depression due to a dissolved solute.

Level

General high school chemistry

Expected Student Background

The major prerequisite for this activity is the ability to read and follow instructions, make and record observations, plot graphs, and safely handle hazardous materials.

Related concepts necessary for understanding this exercise include: atomic structure, formula writing, equations, and concentration calculations which are a direct part of the exercise. Students should know and understand the concept of colligative properties.

Time

The laboratory can be completed in two 45-min periods with *Part I* completed on the first day and *Part II* on another day *without dismantling the apparatus.* Teacher preparation time is about one hour. A great timesaving step is to supply the hot water baths at the proper temperature. The hot water from the tap may be hot enough, though further heating may be necessary in some schools.

Safety

Read the *Safety Considerations* in the *Student Version*. Students should not heat the test-tube of lauric acid directly in the flame as this can cause it to vaporize. It is also flammable. Care should be exercised in using the thermometer, particularly if mercury thermometers are used. Have sulfur or other mercury clean-up materials ready and immediately available in case of accidental breakage if mercury thermometers are used.

Materials (For 24 students working in pairs)

At least one balance with 0.1 g sensitivity should be available to the class.

Nonconsumables

- 24 Ring stands with iron rings and wire gauze
- 24 Thermometer clamps
- 12 Test-tube clamps
- 24 Beakers, 600-mL
- 24 Thermometers
- 12 Burners
- 12 Test-tubes, 25 x 200-mm
- 24 Copper heavy gauge wires with loops for stirring

Consumables

- Lauric acid, 300 g
- Benzoic acid, 30 g
- Paper towels

Advance Preparation

It is important that you do the laboratory yourself beforehand. Time may be saved by pre-filling the test-tubes with about 25 g of lauric acid and having the hot and warm water baths ready. *[NOTE: The mass of the lauric acid in the test-tubes must be determined to the nearest gram.]*

Pre-Laboratory Discussion

Make sure students understand the safety instructions. It is important that you demonstrate the proper placement of the thermometer, clamp and stirrer. Caution students not to attempt to remove or even move the thermometer or stirrer from the solidified lauric acid. A brief discussion of the reason for freezing point depression of solutions as explained in *Answers to Implication and Applications* (Question 4) would be quite useful, too.

Teacher-Student Interaction

One of the most common problems encountered in this procedure is the fact that the molten acid will warm the 30 °C water so that it no longer cools, or the reverse, that warming of the solid acid will by hampered by the warm water cooling off. If you have sufficient water in the beakers, this problem is avoided. During the session, make sure students do not use the thermometers as stirrers, and that they do not attempt to remove the thermometers after the lauric acid has solidified.

Anticipated Student Results

1. The melting point of lauric acid is 44 °C.
2. The average freezing point depression observed is around 2 °C. This translates into a K_f of about 2.5 °C/molal.

Answers to Data Analysis

1. The graphs should consist of a downward slope followed by a flat plateau and then another downward slope after the lauric acid has solidified. In the case of the solution of benzoic acid in lauric acid, the plateau following the initial downward slope is not perfectly level (as discussed in *Answers to Implications and Applications,* Question 4).
2. The freezing point of lauric acid is 44 °C, based on the plateau observed at this temperature on the graph.
3. The freezing point of the lauric acid/benzoic acid mixture is 42 °C, based on the start of the plateau observed on the graph at this temperature.
4. Moles of benzoic acid used: $\frac{2.5\text{ g}}{122\text{ g mol}^{-1}} = 0.020$ mol benzoic acid
5. Molality of the solution: $\frac{0.020\text{ mol benzoic acid}}{0.025\text{ kg lauric acid}} = 0.80\ m$
6. Freezing point depression: 44 °C - 42 °C = 2 °C.
7. $K_f = \frac{2\ °\text{C}}{0.80\,m} = 2.5\ °\text{C}/m$

Answers to Implications and Applications

1. The freezing point and the melting point of all pure substances are the same. (Naturally, superheating or supercooling can provide apparent differences.) There cannot be a "forbidden region" where a substance exists as neither a liquid nor a solid.
2. As we cool the liquid lauric acid, the energy we remove initially comes from the slowing down of the molecules as the temperature drops. The plateau is the point where the energy is lost in the phase change. The lauric acid molecules are forming a solid (crystallizing) and releasing energy as they do so. During the phase change no temperature changes are observed. As we cool the solid lauric acid further the energy we remove comes from the slowing down of the molecules of the solid and the temperature continues to drop.
3. This is the point of the phase change; *i.e.*, the molecules become an ordered solid. Although the kinetic energy of the molecules remains constant, as the liquid solidifies, the potential energy decreases while the molecules become part of an ordered state. (In the opposite direction, when melting occurs, the added energy is absorbed as potential energy as the molecules move into the liquid phase.)
4. The graph of the pure lauric acid shows an abrupt change to a level plateau at the freezing point. In the solution, the lauric acid solidifies, though the molecules of benzoic acid partially disrupt the process and cause solidification to happen at a lower temperature. When the lauric acid solidifies, the benzoic acid tends to stay in the liquid phase and further decrease the freezing point somewhat (*i.e.*, the higher molality of the remaining liquid phase has an even lower freezing point). Thus the plateau is not flat in the solution cooling curve.
5. The lauric acid is crystallizing into an ordered solid state as in the case of the pure acid, though the benzoic acid eventually also precipitates when its solubility in lauric acid is exceeded.

Extensions

1. Repeat the procedure with a larger amount (*e.g.*, 5.0 g) of benzoic acid.
2. Repeat the procedure with a soluble ionic solute to see the particle relationship that exists. A feasible system is water and $CaCl_2$ if you can obtain dry ice from the grocery store to use as a coolant. A dry ice/acetone slurry should be an efficient cooling bath.

Assessing Laboratory Learning

1. Laboratory Practical. A laboratory exercise requiring the determination of the freezing point depression of some other material may be used as assessment. Water is one such solvent.
2. Written Examination. Questions about the molecular level changes and questions requiring calculation of freezing point depression may be used.

Reference

This activity is an extension of a phase-transition experiment by Wilbraham *et al.* (1987).

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Types of Solutions

Purpose

To demonstrate types of solutions.

Safety

See individual demonstration cautions and use generally accepted procedures for safety as outlined in the safety section of the *ChemSource* materials.

Procedures

Gas in gas: Gently warm I_2 crystals in a beaker over a water bath to observe I_2 vapor in the air. The beaker should be covered with a watch glass or evaporating dish. *(CAUTION: Iodine vapors are corrosive to eyes, skin, respiratory tract, and other materials.)*

Gas in liquid: Deaerate well-aerated water by warming (bubbles of air should be apparent) or uncap soda pop bottle previously shaken lightly (CO_2 will be evolved).

Gas in solid: Previously freeze well-aereated water. In class, melt the ice and warm further to deaerate. *(NOTE: Bubbles in the ice are not dissolved gases.)*

Liquid in liquid: Mix ethanol with water. *(CAUTION: Ethanol dehydrates the skin. Use with care. No flames should be present. Avoid breathing solvent vapors.)* The sum of the two volumes is not the same as the volume of the solution and can be shown by careful volume measurement. (100 mL ethanol + 100 mL water ≈ 190 mL solution. The solute fills in the void volume (holes) of the solvent.)

Liquid in solid: The substance in a disposable diaper can dissolve water. The powder can be removed by cutting the diaper into 10 cm^2 pieces. Place the pieces in a plastic bag, close the bag, and shake out the powder. Sprinkle about one-half teaspoon of the powder onto one-half cup water. Observe. (A gel-like solid should form.) Practice this since the amount of water that can be absorbed may vary with the brand of diaper. (This demonstration is adapted from the *Polymer* module. See *Demonstration 3* in the *Polymer* module for another example of this type of solution.)

Solids in liquids (best known): Examples include solutions of sugar or salt (NaCl) in water. Anhydrous $CuSO_4$ (white) dissolved in water (colorless) gives a blue solution. I_2 in different solvents gives different colors (violet in cyclohexane, yellow in ethyl acetate, brown in water).

Solids in solids: Using a soldering iron compare the ease of melting of a typical Sn-Pb alloy (solder) *vs.* Sn metal and/or Pb metal. *(CAUTION: Beware of metal vapors from the solders. Alternatively, compare the properties of brass [Cu-Zn alloy]* vs. *the metallic elements Cu and Zn.)*

Demonstration 2: Colloids, The Tyndall Effect and More

Introduction

The Tyndall effect is usually given as a definitive test to distinguish between a true solution and a colloid. The Tyndall effect involves the scattering of a beam of light as the light passes through a medium having particles of colloidal size. Since particles such as molecules of sugar or sodium ions or chloride ions in solution are too small to scatter light, a beam of light passing through such a solution is not scattered. However, the protein molecules in milk are of colloidal size and consequently a drop of milk mixed into water will cause a light beam traversing the solution to be scattered. To demonstrate the difference, the two systems described above are usually employed. The milk in water changes color as more milk is added (bluish to yellow to red). However, a single system where the particles go from "solution size" to "colloidal size" provides a more dramatic demonstration and as the particles continue to grow additional optical effects may be demonstrated. One such system is the production of sulfur by the reaction between sodium thiosulfate and sulfuric acid.

$$Na_2S_2O_3(aq) + H_2SO_4(aq) \rightarrow S(s) + SO_2(g) + H_2O(l) + Na_2SO_4(aq)$$

Here, the particles of sulfur grow from solution size to colloidal size and finally begin to precipitate. As this phenomenon occurs, it is possible to demonstrate the Tyndall effect and to examine some of the characteristics of the scattered and transmitted light. These characteristics may then be related to other phenomena, such as red sunsets.

Safety

Some students are very allergic to sulfur dioxide (SO_2), which is generated during this demonstration. You should offer a warning and excuse students if they have a known sensitivity to sulfur dioxide (or to sulfites in food, which may signal such a sensitivity).

Option A

Materials

Square or rectangular battery jar or a small fish tank
Stirring rod (long glass rod to stir contents of the battery jar)
Parallel beam light source or a flashlight
Piece of frosted glass or a sheet of white paper stapled to a frame
Ring stand and clamp to hold the glass or paper in a vertical position
Preweighed sample of $Na_2S_2O_3$ or $Na_2S_2O_3 \cdot 5H_2O$ to make a 0.01 *m* solution when dissolved in the water in the battery jar (1.6 g $Na_2S_2O_3$ per L water or 2.5 g $Na_2S_2O_3 \cdot 5H_2O$ per L water)
10 mL Concentrated sulfuric acid, H_2SO_4 per L of water in the battery jar

Optional:

One or more pieces (sheets) Polaroid material
Cardboard silhouette of a flying duck (about the size of half the diameter of the parallel beam) and hung like a mobile from a fine piece of string
Recording of "Canadian Sunset" and appropriate player

Directions

The demonstration should be conducted in a well ventilated area because of the production of $SO_2(g)$ whose odor can be mildly detected while the demonstration is in progress.

Set up the light source, battery jar, and frosted glass as illustrated in Figure 2. The battery jar should be filled with sufficient depth of water so that the entire diameter of the light beam passes through the liquid. (The amount of water should be predetermined in order to have a preweighed sample of sodium thiosulfate and a premeasured volume of concentrated sulfuric acid prepared.) Also, the water should be placed into the battery jar a few minutes prior to beginning the demonstration to permit air bubbles to escape.

Turn on the light source and darken the room. If you will be using the Polaroid sheets, pass these out to students seated directly in front of the demonstration. Ask students whether they can see the beam traversing the water. If using the Polaroid sheets, ask whether rotating them has any effect. *[They should not be able to see the beam but the white round disk of the transmitted beam hitting the frosted glass should be visible. Nothing different should be seen using the Polaroid sheet.]* Point out that the transmitted beam striking the frosted glass is colorless (white). Add the preweighed sample of sodium thiosulfate to the water and stir until it is dissolved. Wait for any bubbles to leave the solution. Repeat the above question(s). *[The reply should be the same.]* Carefully and with stirring, add the premeasured concentrated sulfuric acid to the sodium thiosulfate solution. Wait about 15 seconds for bubbles to clear and repeat the above question(s). *[The reply should be the same.]* Ask if there is any evidence of a chemical reaction? *[There should be no such evidence.]* Ask students to tell you when they notice any change. *[Using the concentrations recommended and a solution temperature of about 20 °C, it should take about two minutes for the Tyndall beam to start to appear and another two minutes for the particles to become large enough and concentrated enough for the percent of transmitted light to go to zero. The reaction rate appears to be first order in thiosulfate concentration so if you want to slow things down, decrease the concentration of sodium thiosulfate.]* As soon as the Tyndall beam becomes visible ask students to compare the color of the scattered light to that of the transmitted light striking the screen. Also ask students using the Polaroid sheets whether rotating them has any effect on the intensity of the scattered light. *[The scattered beam is bluish, whereas the transmitted light starts to turn yellow. As the Tyndall beam becomes more visible the scattered beam loses a little of its bluish color as the green and yellow portion of the spectrum is scattered and the transmitted light striking the screen goes from yellow to red. This observation occurs because the wavelengths of the light scattered are directly proportional to the size of the particle scattering the light. Also, the scattered light is polarized.]* If you have the silhouette of a "flying duck" you can add a dramatic closing touch to the demonstration. Start the recording of "Canadian Sunset" and hang the silhouette over the frosted glass so that it intercepts the disk of the reddening transmitted beam before the beam strikes the glass.

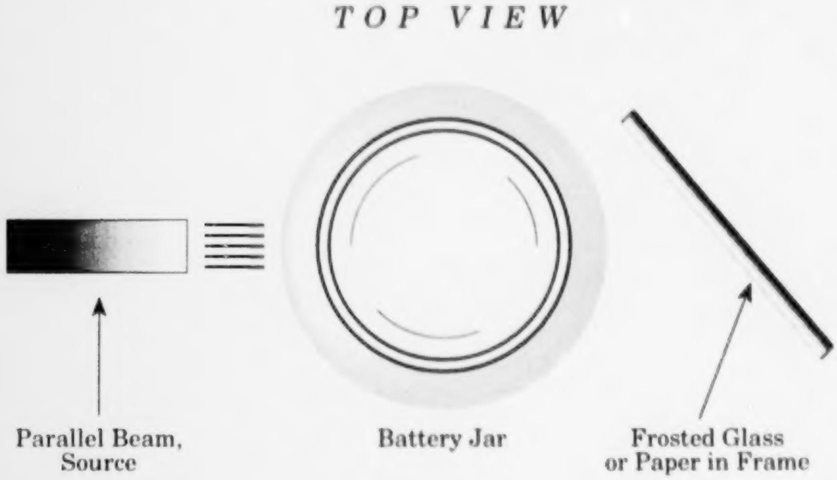


Figure 2. Apparatus for Tyndall Effect demonstration.

Dispose of the battery jar solution as soon as the demonstration is ended to minimize the amount of SO_2 entering the atmosphere.

After summarizing for students that the Tyndall beam was not noted until the particle size had become that of a colloid and that no beam had been visible when there was a true solution, you might want to extend the discussion and demonstration to other areas. If the day happens to be sunny with a relatively clear sky you might ask students to use the Polaroid sheets to observe the light coming from the blue sky at right angles to the direction of the sun. *(CAUTION: Make sure students do not look toward the sun since this is very dangerous.) [The scattered blue light is polarized.]*

Questions

You may follow the demonstration with the following questions:

1. Why does the sun appear exceptionally red when it sets behind a city? *[The blue end of the visible spectrum is scattered by the dust and aerosols in the air over the city and the red end of the spectrum is transmitted.]*
2. Why are fog lights usually yellow and not white? *[The yellow light being of relatively long wavelength is transmitted through the colloidal size fog particles thereby permitting the driver to see, whereas the blue component of white light is scattered back to the driver thereby obscuring vision.]*
3. Why not use red fog lights? *[Red light is usually a signal of danger and approaching motorists may interpret it as such. Also, the human eye is much more sensitive to yellow light than red light.]*
4. Why do things look clearer through rose (pink) colored glasses? *[The pink glass filters out blue light that is scattered by aerosols in the environment. Since this blue light does not reach our eyes, the clarity of objects is increased.]*

Option B

Materials

Overhead projector
Medicine dropper
Beaker of water
Milk, 10 mL

Directions

A simpler and quicker demonstration involves a beaker of water on the stage of an overhead projector and a small amount of milk in a medicine dropper. As the milk is slowly added to the water, the color of the water becomes bluish, while the circle of light thrown up onto the screen gradually becomes yellow then orange and then red. The questions suggested above (*Option A*) are also appropriate for this option.

Demonstration 3: Paper Chromatography

Introduction

Chromatography is used to separate the components of a mixture, operating on the basis of selective adsorption. The solvent moves through the porous material called the adsorbent. The various parts of the mixture are attracted to the adsorbent differently. The more strongly attracted parts do not move as far as the more weakly attracted parts. It is possible to calculate an **R_f value**, which is defined as the distance traveled by the solute divided by the distance traveled by the solvent. The purpose of this demonstration is to separate the pigments in a food color by paper chromatography.

Safety

Proper laboratory safety procedures must be followed. Alcohols are flammable; no open flames are allowed. Any substances spilled should be immediately washed off with large amounts of water.

Materials

2 Petri dishes (less than 11 cm diameter)
Capillary tube (thin-walled, open-ended melting point tube or Pasteur pipette)
Filter paper, 11 cm
Ethanol, 95% or absolute (Solvent 1)
0.1% Sodium chloride, NaCl (0.1 g NaCl or table salt per 100 mL water; Solvent 2)
Food coloring—Green typically works best, but try more than one color
Scissors
Distilled water

Procedure

1. Cut two pieces of filter paper as shown in Figure 3. Fold the tongue down so that it will just touch the surface of the Petri dish.
2. Using the capillary tube to transfer the food dye to the paper, make a small spot behind the folded tongue with one of the samples of food coloring. When the spot is dry, touch the capillary tube to the paper again to increase the amount of pigment on the paper. Do not increase the size of the spot; just try to make it more concentrated.
3. Prepare the other sample in the same way.
4. Place enough of one of the solvents to cover the bottom of one of the Petri dishes. Repeat with the other solvent in another Petri dish.
5. Place one piece of filter paper over each dish with the tongue just touching the solvent.

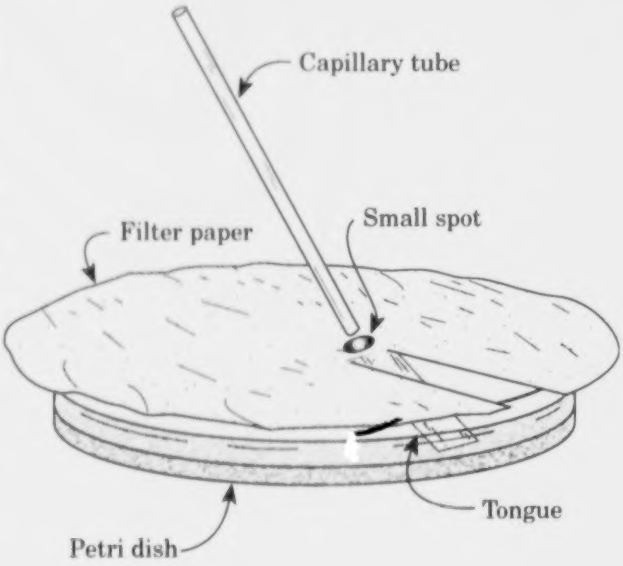


Figure 3. Set-up for paper chromatography demonstration.

6. After the solvent is well absorbed (*i.e.*, traveled along) onto the paper without completely moving to the edge, remove the filter paper and mark the outer edge of the solvent line using a pencil. Do not allow the solvent front to reach the edge of the filter paper. When the papers have dried, mark the edges of each pigment using a pencil.
7. Calculate the R_f value for each spot in each solvent.

$$R_f = \frac{\text{distance traveled by solute}}{\text{distance traveled by solvent}}$$

(NOTE: R_f values are always ≤ 1, and that a smaller R_f value corresponds to a more tightly held solute.)

8. If time permits redo with other colors.

Data Analysis

1. How far did each solvent travel?
2. How many pigments were observed?
3. How far did each pigment travel from the original spot?
4. Calculate the R_f value of each major pigment.
5. Which color had the largest R_f value?
6. Was the same order of elution (as determined by the order of the colors observed) for each solvent?

Extensions

1. Many inks and dyes can be separated in this way (see *Demonstration 4*).
2. Likewise you can try many other solvents, *e.g.*, nonpolar organic solvents. Mixed solvents (*e.g.*, aqueous ammonia and ethanol) can also be used.
3. This demonstration can also be developed into a full laboratory experiment.
4. Chromatographic paper used in chromatographic developing tanks can also be used but is unnecessary. Strips or cylinders of rectangular chromatographic paper in small glass jars with covers or beakers with Petri dish covers can also be used. The one directional movement allows the development of several spots at once.
5. For a detailed experiment using paper chromatography with food dyes, see Markow, P. G. (1988). *J. Chem. Ed*, *65*, 899-900. This paper also references six other versions of this experiment or demonstration for teaching the basics of chromatography.

Demonstration 4: Separation of the Components of a Mixture by Column Chromatography

Introduction

The purpose of this activity is to separate the components of a mixture by column chromatography, a very common method for separating a mixture. It is based on the fact that different materials are attracted to adsorbents with varying strengths. These attractions are based on the types of forces of attraction; very polar solutes are strongly attracted to some column media while nonpolar solutes are not. The solvent also plays an important role in the proper operation of the column. If the solvent attracts the solutes too strongly, all of the solutes travel close to the solvent front (high but similar R_f values) and good separation is not obtained.

Safety

Proper laboratory safety procedures must be followed. Any substances spilled should be immediately washed off with large amounts of water.

Materials

Support stand and clamps
Chromatography column*
Cellulose packing*
Glass wool*
Dyes to be separated*
Distilled water should also be available.
*These materials are contained in a kit available from Sargent-Welch; stock number S-18818-10. The refill materials are also available; stock number S-18818-20.

Procedure

It is important that you do the demonstration prior to doing it before a group. It will require several hours to set up and perform.

1. Support the column with clamps on a support stand. Be sure that the column is vertical and that there is sufficient room to place a collecting flask under it.
2. Place a small quantity of glass wool in the bottom of the column.
3. Close the stopcock and fill the column with distilled water (solvent).
4. While tapping the column (use your thumb or the eraser head on a pencil) sprinkle the cellulose into the column slowly. The purpose of the tapping is to help the packing material settle uniformly in the solvent. It will take about 15-20 min to fill the column.
5. Once the column is filled, place a plug of glass wool on the top. The purpose of this plug is to protect the surface of the packing while you add sample and solvent. Open the stopcock allowing solvent to run out of the column until the top level of the solvent is even with the top of the upper glass wool plug, and then close the stopcock.
6. Dissolve a small amount of the solute in 10-15 mL water; use a dropper to place it carefully onto the column.
7. Allow the solvent to run out until the level of the pigment is even with the top of the upper glass wool plug.
8. Carefully fill the column with solvent, then allow the column to run. *CAUTION: You must maintain the solvent above the top of the glass wool plug atop the column. Do not allow the column to run dry.*
9. Once separation is evident you can show students and explain the system.

Data Analysis

1. What evidence do you have that separation occurred?
2. Which colored substances moved the fastest? The slowest?

Extensions

1. If you continue running the column, and collect the different fractions in separate beakers you can demonstrate how complete separation is attained.
2. Try other samples and solvent systems.

Suggestions for Other Demonstrations

CAUTION: Use appropriate safety guidelines in performing these activities.

Dilution: *CEPUP* (Chemical Education for Public Understanding Program, Lawrence Hall of Science, University of California, Berkeley) has a semiquantitative dilution experiment in which dilution is continued on a multi-well strip or plate until the solution appears colorless.

Supersaturation: *Doing Chemistry* DMEX C22 (Videodisc side C, Frame 20792, and Teacher's Manual, p. C22-48) provides details for the classical sodium acetate hydrate supersaturation demonstration. (See also *The Chemistry of Rocks, Minerals, and Gems* module.)

Distillation: *ChemCom* (Chemistry in the Community) has a distillation activity using ethylene glycol and water that can be done either as a laboratory exercise or as a demonstration.

Conductivity of Solutions: *Doing Chemistry* DMEX C26 (Videodisc Side C, Frame 25883, and Teacher's Manual, p. C26-62) provides details for a small-scale electrical conductivity demonstration. However, a simpler small-scale electrical conductivity meter can be constructed from a battery and a low-voltage light bulb. Either type is safer than using 120-volt electricity as the electrical circuit, although the use of direct current has the disadvantage that electrochemical deposition can ruin the electrodes if used for long time periods (see *Periodicity* module).

Ice Cream: A fun activity is to study colligative properties through the making of ice cream. The students should undoubtedly perform the activity in the kitchens of the school's Home Economics facilities or at home rather than in the chemistry laboratory. Use a thermometer to show the depressed freezing point of the ice-salt bath, *etc*.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. Distinguish between the solute and the solvent in each solution in the laboratory and demonstration activities of this unit. *[NOTE: The selection is arbitrary in some cases.]*
2. Why do colloidal particles remain dispersed throughout the solvent instead of falling to the bottom of the container? *[See* Content in a Nutshell.*]*
3. Why is molarity (*M*) a convenient unit for solution reactions? *[Solutions used in reactions are usually measured by volume. The stoichiometry of reactions is based on moles. Thus, moles of solute per liter of solution is the most convenient unit.]*
4. Why are so many different concentration units used? Consider the types of situations when each might be preferred. *[Molarity is discussed in the answer to Question 3. Molality is useful for colligative property studies to keep the amount of liquid solvent constant; mole fraction is useful for colligative properties of gases; % by mass for components that are easily weighed; % mass to volume for solids in liquids; % by volume and volume:volume for liquids in liquids; and ppm for trace amounts of solids in liquid solvents.]*
5. Differentiate the properties that would allow the use of distillation for the separation of solution components from those that would allow the use of crystallization. *[Distillation would be necessary if both components were liquids, whereas crystallization is satisfactory for solid components.]*
6. Why are the colligative properties of aqueous solutions affected more by a given number of moles of an electrolyte than by the same number of moles of a nonelectrolyte? *[Colligative properties depend on the number of particles and electrolytes have more particles per mole as a result of ionization or dissociation.]*

Counterintuitive Examples and Discrepant Events

1. When a dog urinates on grass, urea (a fertilizer) is added to the soil in high concentration. Why does the grass die? *[The osmotic imbalance dehydrates the cells in the grass, and it dies.]*
2. Deep-sea divers use helium/oxygen gas solutions rather than air for breathing because of the problem (the bends) that occurs when nitrogen is used. Why? *[Because of the high solubility of nitrogen in water at higher pressures. Helium is much less soluble.]*

TIPS FOR THE TEACHER

Language of Chemistry

Concentration Terms

molarity (molar concentration) (M) moles of solute per liter of solution.

molality (molal concentration) (m) moles of solute per 1000 grams (1 kg) of solvent.

mole fraction (X_1) moles of component 1 divided by the total moles of all of the solution components (including the solvent).

percent by mass mass of solute divided by the total mass of all of the solution components (including the solvent).

percent mass to volume mass of the solute relative to the volume of solvent.

percent by volume volume of the solute divided by the total volume of the solution.

volume:volume the volume ratios of the solution components.

parts per million (ppm) number of mass units of the solute of interest relative to one million mass units of solution; *e.g.*, micrograms per gram or milligrams per kilogram.

NOTE: The last five terms are not on a particle or mole basis for either solute or solvent!

- Molality is moles per 1000 *g* of *solvent* whereas molarity is moles per 1000 *mL* of *solution*.
- Mole fraction is the only concentration unit that is consistent from solvent to solvent.

solution types

A **solution** is a homogeneous mixture. Although solutions are often thought of in terms of liquid solvents, solutions can be solids in solids (alloys, such as brass—zinc in copper), liquids in solids (mercury in zinc—referred to as zinc amalgam), gases in solids (hydrogen in palladium), solids in liquids (salts in water), liquids in liquids (ethylene glycol or ethanol in water), gases in liquids (air in water), and gases in gases (air). Although liquids and solids in gases are theoretically possible, if the particles are small enough to provide a *homogeneous* material in the gas, the material is effectively a gas itself.

colloids

Colloidal particles, which must be large enough to diffract or scatter light, can be either large single molecules (egg albumin, a protein of molar mass about 43,000, or hemoglobin, the oxygen-carrier in your blood of molar mass about 68,000) or clusters of normally insoluble materials (colloidal gold in water or dust in air).

Pattern Recognition

One of the most important aspects of this unit is solving solution concentration problems. Inasmuch as most textbooks have many solution concentration problems, none have been included other than those included with the laboratory and demonstration activities. Instilling the molarity concept is most important for reaction stoichiometry; however, selected problems using the other common concentration units (see *Language of Chemistry*) should also be included.

Common Student Misconceptions

1. **"The solution and melting processes are the same."**

   This misconception is illustrated by the conversation:

   Warning: "Are you going out? It's raining very hard!"
   Response: "Don't worry, I won't *melt*."

2. **"It doesn't matter whether a liter of solution or 1000 g of solvent are used in calculations."**

   Because molarity and molality are very similar in dilute aqueous solutions, students often assume that the above statement is true. To avoid this confusion, use solvents with densities other than approximately 1.0 g/mL and check aqueous solution calculations very carefully to insure that the right units are being used.

Decision Making

Challenge students to consider the practical methods for desalination of sea water. Processes that have been considered include reverse osmosis, ion exchange, freezing, nuclear-powered distillation, and solar vaporization and recondensation. Have them evaluate the relative costs, volume of water per unit of space required, and other factors that they feel are important to finding the best solution to the long term water shortage problem that exists in many parts of the world.

HISTORY: ON THE HUMAN SIDE

1. When Arrhenius developed the concept of electrolytes in solution, he was still a student. The idea that ions were moving about in water was considered so preposterous by his doctoral committee that he just barely passed his doctoral examination (see *Acids and Bases* module).
2. Solutions have played an important part in human history for many centuries. It is believed that wealthy Roman citizens met early deaths because of lead poisoning from the lead pipes they used. Other Romans, lacking the luxury of piped water, were not affected by the problem.
3. Shortly after the introduction of detergents containing sulfonate groups, lakes and streams developed huge amounts of foam. The synthetic detergents did not break down. A switch to phosphate detergents solved the foam problem, but led to eutrophication (see definition in *Links and Connection* section) of lakes and streams instead. The reformulation with minimal phosphate content has minimized that situation considerably though further research is continuing to develop even better detergents.

HUMOR: ON THE FUN SIDE

1. "Mathematicians have problems, but chemists have solutions." [Students could make bumper stickers.]
2. $\frac{2\ \mathrm{NaCl(aq)}}{\mathrm{C_7}}$ *[Saline, saline, over the seven C's]*
3. Then there are the witches from William Shakespeare's *Macbeth*:

   "Double, double, toil and trouble,
   Fire burn and cauldron bubble."

   Even the great bard of the English language was into making solutions!

4. A modern young man thought he'd boobed
   When the milk he had ordered came tubed
   Instead of "One liter"
   The package read (neater)
   "Contents are one dm cubed."

   (CHEM 13 NEWS, September 1978, p. 25)

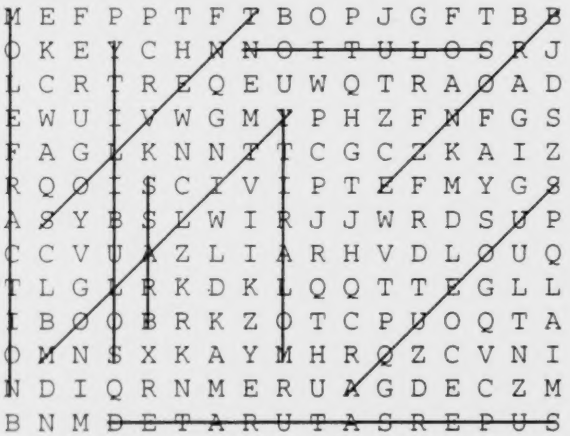
5. a. **Question**: How do you make antifreeze?
      Answer: Hide her scarf and mittens.

   b. Solute is what you give a general.

   (CHEM 13 NEWS, December 1976, p. 1123)

6. **Another message on a bumper sticker**: Chemists have solutions...

7. Word Search (see *Appendix* for master copy)



M E F P P T F Z B O P J G F T B B
O K E Y C H N ~~N O I T U L O S~~ R J
L C R T R E Q E U W Q T R A O A D
E W U I V W G M Y P H Z F N F G S
F A G L K N N T T C G C Z K A I Z
R Q O I S C I V I P T E F M Y G S
A S Y B S L W I R J J W R D S U P
C C V U A Z L I A R H V D L O U Q
T L G L R K D K L Q Q T T E G L L
I B O O B R K Z O T C P U O Q T A
O M N S X K A Y M H R Q Z C V N I
N D I Q R N M E R U A G D E C Z M
B N M ~~D E T A R U T A S R E P U S~~

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Moles of solute per liter of solution.
2. Homogeneous mixture of two or more substances that are at the smallest levels of their states of subdivision.
3. Solution of copper and zinc.
4. Solution of copper and tin.
5. Moles of solute per 1000 grams of solvent.
6. Type of unstable solution that has a higher concentration of solute than normal at a given temperature.
7. Extent to which a substance dissolves in a fixed volume or mass of a solvent at a given temperature.
8. Adjective describing a solution in which water is the solvent.
9. That solution component into which the solutes are considered to have dissolved; the component that is present as the continuous phase.
10. Ratio of moles of one component divided by the total moles of all solution components, including the solvent. (two words)

Answers: 1. MOLARITY 2. SOLUTION 3. BRASS 4. BRONZE 5. MOLALITY 6. SUPERSATURATED 7. SOLUBILITY 8. AQUEOUS 9. SOLVENT 10. MOLE FRACTION

8. See cartoons at end of module.

MEDIA

1. The World of Chemistry videotape "Number 10: Water" has some good materials on solutions. It shows water dissolving a salt, discusses the polarity of water. discusses miscibility–immiscibility, water-soluble polymers, *etc*. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    - a. *Animated Demonstrations*, by Philip Pavlik. Vol. V B, No. 2, for IBM PS/2. PC-compatible computers.
    - b. *The Acid-Base Package*, by Richard Ramette. Vol. II B. No. 2, for IBM PS/2. PC-compatible computers.
    - c. *Acid-Base Package*, by Richard Ramette and Jon L. Holmes. Vol. IV C. No. 2, for the Apple Macintosh.
    - d. *Equilibrium Calculator*, by Robert D. Allendoerfer. Vol. VI B, No. 1. for IBM PS/2. PC-compatible computers.
    - e. *Equilibrium Calculator*, by Robert D. Allendoerfer, Vol. I D. No. 1, for Windows running on IBM PS/2, PC-compatible computers.
3. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
    - a. For the Apple II computer: AP 502, AP 506.
    - b. For the Apple Macintosh: MC 50 1.
    - c. For IBM PCs and PC-compatibles: PC 2501
4. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    - a. "Changes of State in Bromine," "The Hydrologic Cycle," "Acid Rain and Limestone" and "Lime Water and Carbon Dioxide," four chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.
    - b. "A Solution with Water," "Alcohol Dissolves in Water," and "Water and Oil," three chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

5. Additional software available from QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
    a. *J & S Solutions* (J & S; Apple-II and IBM-PC) Naming of compounds, calculating ion concentration, molarity problems, calorie problems, and the formation of precipitates.
    b. *Solutions and Near Solutions* College Board Achievement Test in Chemistry Series (Intellectual Software, Queue; Apple-II, IBM-PC, and Macintosh) Properties of solutions, molarity, molality, normality, colloidal suspensions, Brownian motion, Tyndall effect, hydrates, and efflorescent and deliquescent substances.
    c. *Solution Equilibrium, Colligative Properties, and Acids and Bases.* College Board Achievement Test in Chemistry Series (Intellectual Software, Queue; Apple-II, IBM-PC, and Macintosh) Vapor pressure depression, boiling point elevation, freezing point depression, osmotic pressure, and other topics.
    d. *FM Solutions* (Focus Media; Apple-II) Types of solutions, solubility factors and curves, and solution problems.
    e. *SEI Solutions* (Sliwa Enterprises; Apple-II and IBM-PC) Types of solutions and chemical equilibria.
    f. *Topic 11: Solutions.* General Chemistry 1A Series for AP or College Chemistry (Knowledge Factory; Apple-II) Solution terminology, concentration, colligative properties, electrolytes and non-electrolytes, separation methods, and diffusion.
6. *Solutions: Ionic and Molecular,* 23 min. video available from Coronet/NTI Film and Video, 108 Wilmot Road, Deerfield, IL 60015; (800) 621-2131; (708) 940-3640 (FAX).

T6

Links/Connections

WITHIN CHEMISTRY

Solutions and solution concentrations are very important in solubility and precipitation, chemical reactions, stoichiometry, chemical equilibria, acids and bases, electrochemistry, separations, qualitative analysis, and rates of reactions. See the respective *SourceBook* modules.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Biological sciences require a thorough understanding of aqueous solutions and concentrations. Health-related sciences—as pharmacy, nursing, medicine, and dentistry—make extensive use of solutions. Solutions are also of concern in chemical engineering and related fields. Because both air and natural waters are solutions, environmental scientists are considerably involved with solutions.

TO THE CONTEMPORARY WORLD

Personal

Consider the concentrations of key ingredients of various items in your home. Look at a box of breakfast cereal that is vitamin fortified and note the concentration units that are used. Check other food and beverage items as well. If you have a bottle of tincture of iodine in your medicine chest, note the units of concentration that are used. Check other over-the-counter drug items; *e.g.*, 0.09% saline, 3% hydrogen peroxide. How about shampoos? Cleaning solutions? Adhesives?

Community

Both air and natural water supplies are solutions. Water supplies typically have many ions and polar molecules dissolved in them. Consider the water supply system in your community. Is your water considered a "hard-water" supply? Why is it called hard water? How can you test for the hardness? (See *Biogeochemical Cycles* module.) How is the level (concentration) of chlorine in your water determined? Find out by contacting the appropriate "water chemist" in your community. Are there pollutants in your water supply? How are they measured? Or are they? What about air pollution in your community? Acid rain? More detailed information can be found in the *Chemistry in the Community* (*ChemCom*) units on water and the atmosphere.

Other ecological situations can be considered; *e.g.*, DDT (an organic pesticide once in very wide use), accumulates in the fatty tissues of animals. Determine how this fact fits with the classic expression: "like dissolves like." Also water passing through a landfill or garbage dump dissolves many substances that can eventually contaminate water supplies. These problems involving solutions can provide interesting class discussions, especially when probed to find answers (solutions!) to these problems.

Many workers in your community use solutions in their work, whether they know it or not. Have students think of persons who deal with solutions. A few hints follow: Gasoline has additives in addition to its mixed hydrocarbon content. Antifreeze is an ethylene glycol/water solution with added rust inhibitors. Hair stylists use "permanent" wave solutions and shampoos, which may be either true solutions or colloids. Many food and beverage solutions can be obtained in stores and restaurants. [The brine (aqueous NaCl solution) used in pickles provides a good example of solution control. If the NaCl concentration in the brine is too high, the pickles will shrivel up.] The use of solutions in the medical field was briefly noted above. Divers use a helium/oxygen gas mixture for breathing in order to avoid "the bends." (Helium is not as soluble in water as is nitrogen.)

The World

The oceans provide the most abundant aqueous solution on earth. The seas contain 97.6% of the water on this planet. The water content of the ice caps and glaciers (1.9%), ground water (0.5%), rivers and lakes (0.02%), soil moisture (0.01%), and the atmosphere (0.0001%) is significantly lower. Although we normally consider ocean water as salt water, ions other than Na^+ and Cl^- are also quite prevalent as are dissolved gases and organic materials. A consideration of the components of the ocean water is an interesting exercise. See the article on "Oceans and Seas" in the *Encyclopaedia Britannica* or the *Chemistry of Seawater* module for details.

Naturally, the atmosphere is another important solution composed of nitrogen, oxygen, argon, carbon dioxide, water vapor, smaller amounts of sulfur oxides and nitrogen oxides, *etc.* Colloidal dust particles exist in both water and air supplies.

Another important solution example is the *eutrophication* of fresh-water lakes and streams, *i.e.*, the overgrowth of algae due to excessive nutrients (nitrates and phosphates, in particular) in the water from fertilizer runoff from homes, farms, golf courses, *etc.* and from detergents (phosphates).

References

Module developed by R. D. Archer, W. G. Cumming, and A. M. Rennert, the New England team.

Brooks, D. W. (Producer). (1989). Doing chemistry [Videodiscs, computer program, and supporting written materials]. Washington, DC: American Chemical Society.

Lippincott, W. T. (Principal Investigator) and American Chemical Society team. (1988). *ChemCom: Chemistry in the community.* Dubuque, IA: Kendall/Hunt.

Concept of solutions is discussed in "Investigating the Cause of the Fish Kill."

Markow, P. G. (1988). The ideal solvent for paper chromatography of food dyes. *Journal of Chemical Education*, *65*, 899-900.

The New Encyclopædia Britannica. (1986). Chicago, IL: Encyclopedia Britannica.

The articles on "Solutions" and on "Oceans and Seas" are relevant to this module.

Wilbraham, A. C., Staley, D. C., Simpson, C. J., and Matta, M. S. (1987). *Chemistry laboratory manual.* Menlo Park, CA: Addison Wesley.

Laboratory Activity 2 is an extension of a phase-transition experiment in this manual.

Appendix

- **Transparencies**
  1. Solute Concentration Units
  2. Aqueous Salt Solution
  3. Word Search
- **Humor**

Solute Concentration Units

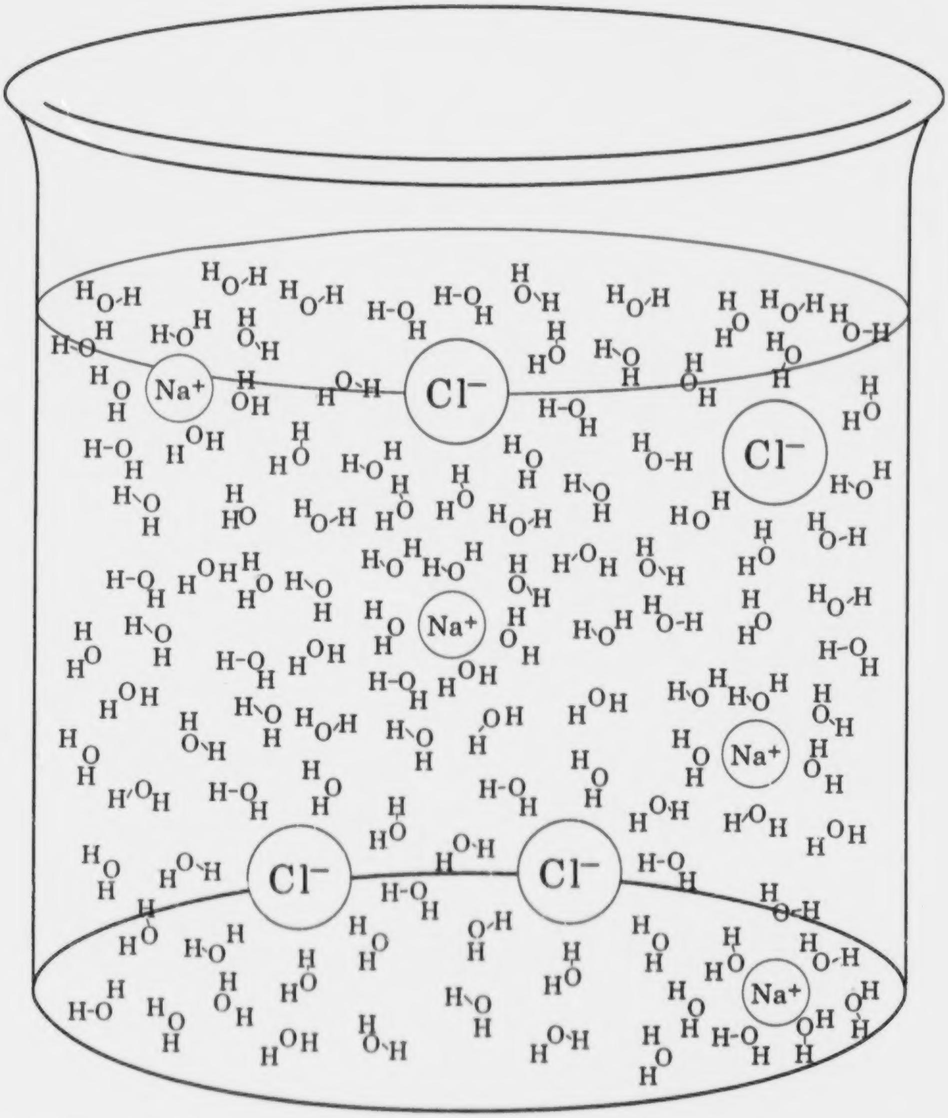
Molarity = Moles of solute/liter of solution

Molality = Moles of solute/1000 g of solvent

Mole fraction = Moles of solute/moles of solution*

*including moles of solvent

Aqueous Salt (NaCl) Solution


Na⁺
Cl⁻
Cl⁻
Na⁺
Na⁺
Cl⁻
Cl⁻
Na⁺

*Note H end of water nearest Cl⁻ and O side of water nearest Na⁺.

Word Search

```
M E F P P T F T B O P J G F T B B
O K E Y C H N N O I T U L O S R J
L C R T R E Q E U W Q T R A O A D
E W U I V W G M Y P H Z F N F G S
F A G L K N N T T C G C Z K A I Z
R Q O I S C I V I P T E F M Y G S
A S Y B S L W I R J J W R D S U P
C C V U A Z L I A R H V D L O U Q
T L G L R K D K L Q Q T T E G L L
I B O O B R K Z O T C P U O Q T A
O M N S X K A Y M H R Q Z C V N I
N D I Q R N M E R U A G D E C Z M
B N M D E T A R U T A S R E P U S
```

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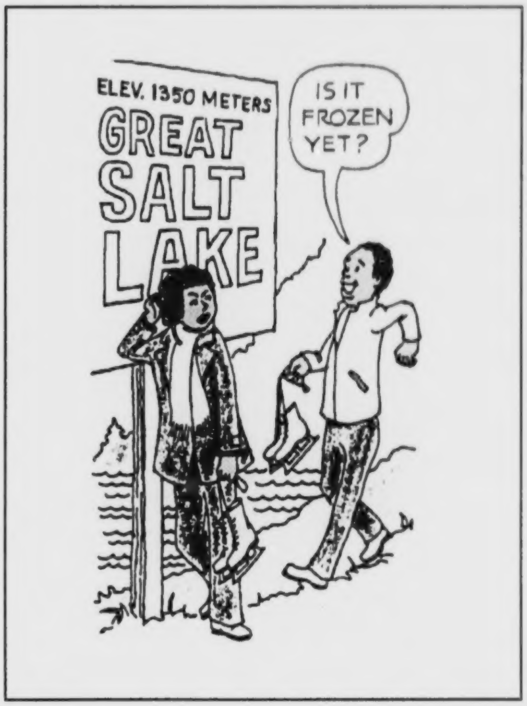
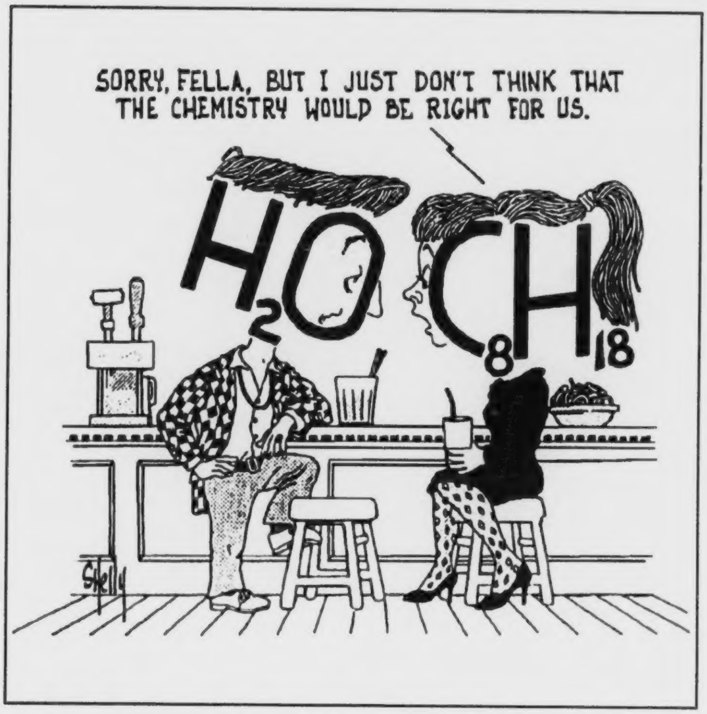
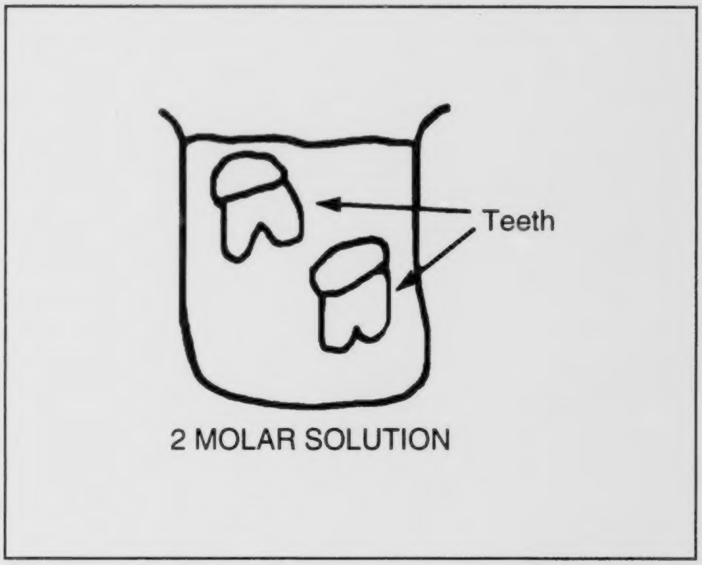
1. Moles of solute per liter of solution.
2. Homogeneous mixture of two or more substances that are at the smallest levels of their states of subdivision.
3. Solution of copper and zinc.
4. Solution of copper and tin.
5. Moles of solute per 1000 grams of solvent.
6. Type of unstable solution that has a higher concentration of solute than normal at a given temperature.
7. Extent to which a substance dissolves in a fixed volume or mass of a solvent at a given temperature.
8. Adjective describing a solution in which water is the solvent.
9. That solution component into which the solutes are considered to have dissolved; the component that is present as the continuous phase.
10. Ratio of moles of one component divided by the total moles of all solution components, including the solvent. (two words)


ELEV. 1350 METERS
GREAT SALT LAKE
IS IT FROZEN YET?

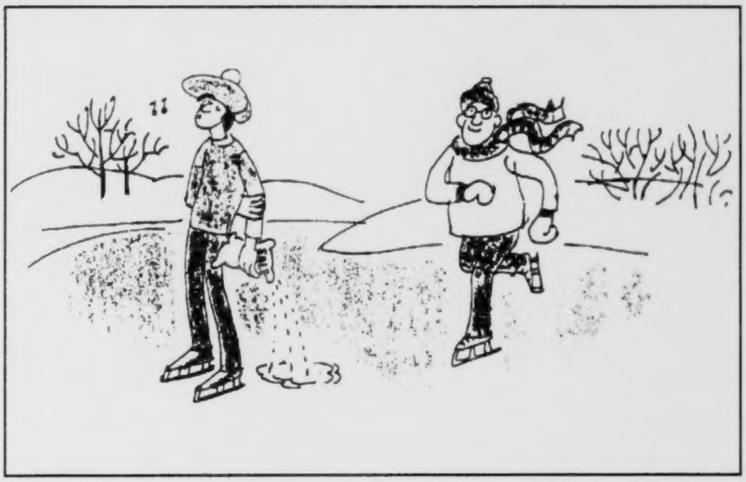
Used by permission of Veritex Publishing Co.


SORRY, FELLA, BUT I JUST DON'T THINK THAT THE CHEMISTRY WOULD BE RIGHT FOR US.
H_2O
C_8H_{18}
Shelly

Used by permission of McGraw-Hill Book Co.


Teeth
2 MOLAR SOLUTION

CHEM 13 NEWS, October 1977, p. 1249. Reprinted with permission.



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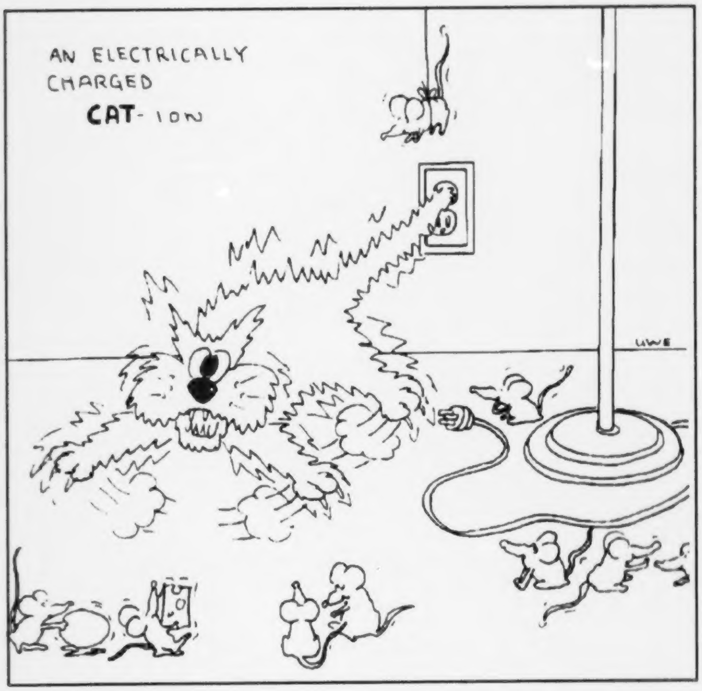
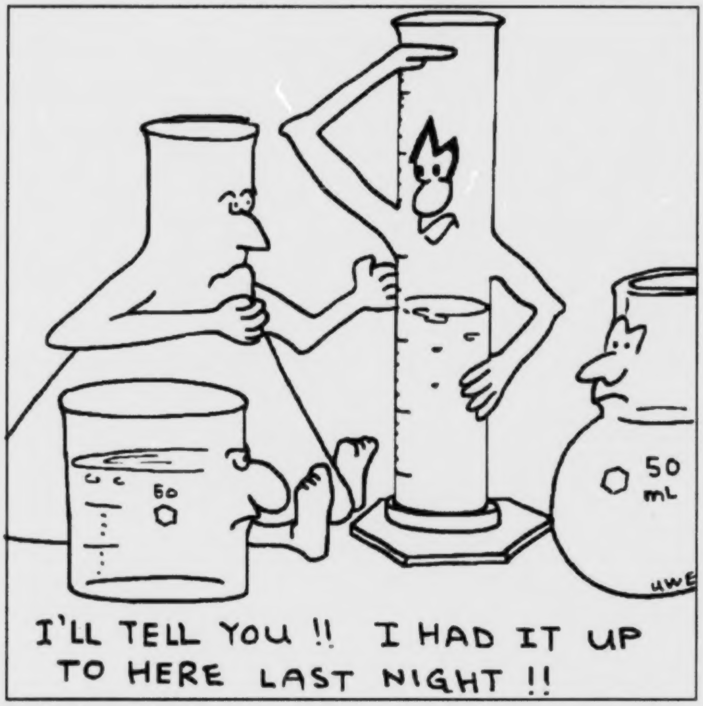
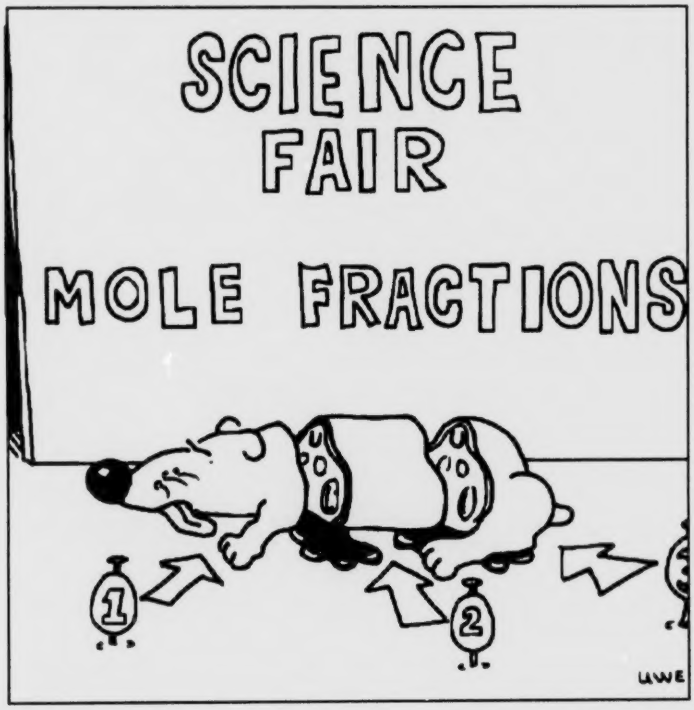

AN ELECTRICALLY
CHARGED
CAT-ION
UWE

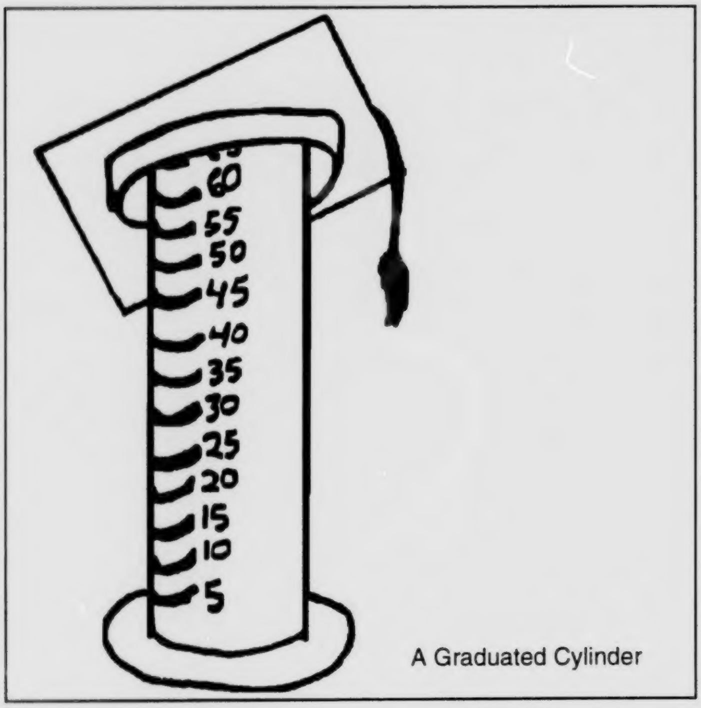
CHEM 13 NEWS, April 1979, p. 5. Reprinted with permission.


60
50 mL
UWE
I'LL TELL YOU !! I HAD IT UP TO HERE LAST NIGHT !!

CHEM 13 NEWS, November 1979, p. 1. Reprinted with permission.


SCIENCE FAIR
MOLE FRACTIONS
1
2
UWE

CHEM 13 NEWS, January 1982, p. 7. Reprinted with permission.


60
55
50
45
40
35
30
25
20
15
10
5

A Graduated Cylinder

CHEM 13 NEWS, March 1981, p. 4. Reprinted with permission.

A SourceBook Module

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STOICHIOMETRY

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Topic Overview

CONTENT IN A NUTSHELL

The big question in chemistry laboratory work is often "how much?" This same question pervades the realm of chemical research and industrial production. Since the work of early scientists such as Lavoisier, one of the first to make careful, quantitative observations about reactions, this question has centered around the understanding of the Law of Conservation of Mass.

With the development of the mole concept, the basis for quantitative evaluation of chemical reactions has been associated with the balanced chemical equation, which derives its validity from the Law of Conservation of Mass. The study of the quantitative relationships in chemical reactions is commonly known as **stoichiometry**, deriving its name from the Greek words *stoicheion* for element and *metrein* for measure.

Since the work of Dalton, quantitative relationships of atoms combining with each other to form compounds of definite ratios has continued to organize chemists' approach to the study of stoichiometry. Chemists today use equation coeffiecients as the basis for quantitative study of chemical change. The equation coefficients represent moles of substances reacting and moles of substances produced (assuming the reaction goes to completion). This is very convenient since the amount of a substance can be expressed in terms of moles.

Frequently chemical reactions involve solutions. Using molarity to express the concentration of a solution facilitates stoichiometric calculations. For gases, the ideal gas law can be used to relate volume to moles. (This law is acceptable for most laboratory-scale reactions but may not be suitable for industrial-scale processes taking place at high pressure.) The important principle is that equation coefficients are the basis for stoichiometric calculations, and the means by which the number of moles involved in the reaction is determined is secondary.

When the concepts of limiting reactant and percentage yield are included, the study of stoichiometry begins to have added significance in industrial and research applications. It is important, however, not to overstate the idea that the limiting reactant "controls" the extent of a reaction. Most reactions encountered in beginning chemistry courses have equilibrium positions displaced toward products. However, all reactions do not "go to completion." In such reactions, the limiting reactant is not the controlling factor; the percent yield may be much less than 100%. (Factors affecting the position of equilibrium are discussed in the *Chemical Equilibrium* module; kinetic factors are considered in the *Rates of Reaction* module.)

PLACE IN THE CURRICULUM

In general and honors courses stoichiometry is discussed early in the course, usually after the mole concept discussion as an application example. For students with weak mathematics skills, stoichiometry may be delayed; however, balancing equations, discussed in this module, may be presented earlier.

CENTRAL CONCEPTS

1. The coefficients of a chemical equation are the basis for interpreting and predicting quantitative relationships in chemical reactions (stoichiometry).
2. Stoichiometric relationships are based on the mole concept.
3. The reactant present in the smallest stoichiometric amount (the limiting reactant) determines the amounts of other reactants consumed and the amount of product formed.
4. The actual yield of product in a reaction may be less than the theoretical yield.

RELATED CONCEPTS

1. Mass, volume
2. Formulas (empirical and molecular)
3. Atoms, molecules, formula units
4. Accuracy and precision
5. Gas behavior (particulate nature of matter, relationship of volume to moles of gas)
6. The mole
7. Chemical equations
8. Reactant, product
9. Percent
10. Conservation of matter
11. Solution concentration (molarity)

RELATED SKILLS

1. Exponential numbers
2. Significant figure manipulation
3. Problem solving strategies (dimensional analysis, *etc.*)
4. Algebraic manipulations
5. Graphing
6. Observation
7. Mass, volume and temperature measurements
8. Use of burner/hot plate

PERFORMANCE OBJECTIVES

After completing a study of stoichiometry, students should be able to:

1. explain how the total mass before reaction is related to the total mass after reaction (whether or not the reaction goes to completion).
2. explain how a balanced chemical equation relates to the law of conservation of mass.
3. balance chemical equations and explain how the coefficients are related to the amount (in moles) of reactants and products.
4. determine the amount (in moles) of any substance in a reaction, given the amount (in moles) of any other substance.
5. determine the mass of any substance in a reaction, given the amount (in moles) of any other substance.
6. determine the mass of any substance in a reaction, given the mass of any other substance.
7. explain what is meant by "limiting reactant."
8. identify the limiting reactant in a reaction and calculate the theoretical yield of product(s) and the quantity of other reactant(s) consumed or unreacted.
9. explain the meaning of percent yield and explain why actual yield is often less than theoretical yield.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Stoichiometry of a Reaction and Limiting Reactant

Introduction

A chemical reaction can be described by an equation. The equation shows the relationship of the moles of reactants to moles of products. Since matter is neither created nor destroyed in chemical reactions (Law of Conservation of Mass), the equation must show equal numbers of atoms of each element in both reactants and products. Such an equation is said to be "balanced."

Stoichiometry is the study of quantity relationships in chemical reactions. The stoichiometry of a reaction is based on the mole. The coefficients in the balanced equation represent the relative numbers of moles of substances that react or are formed in the reaction.

For practical reasons, the quantities of the reactants used in a reaction are generally not in the same mole proportions as the balanced chemical equation. It is important to be able to predict how much product will form under these conditions.

Purpose

To demonstrate that substances react in proportions related to the number of moles of reactants. To show that if the reactants are not present in stoichiometric proportions, the substance present in smallest stoichiometric amount (limiting reactant) will determine the amount of product(s) formed.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Carefully clean up any spills.
3. Exercise caution in lighting the burner. Loose hair should be tied back. Hair ignites easily especially if hair spray is used.
4. Always exercise caution when handling glassware especially if it is hot. Hot and cold glass look the same!

Procedure

1. Weigh accurately a clean, dry 50- or 100-mL beaker. Record the mass. Add to the beaker the assigned quantity (0.2-0.5 g) of iron filings. Reweigh accurately the beaker and contents. Record the mass.
2. Weigh accurately a 150- or 250-mL beaker. Add approximately 3.0 g $CuSO_4 \cdot 5H_2O$. Reweigh accurately and record the mass of beaker and $CuSO_4 \cdot 5H_2O$. Add approximately 50 mL water (preferably distilled water) to the beaker.

   Warm the beaker and stir carefully to dissolve the solid. (Do not boil the solution.) Record the solution's color.

   Alternate procedure: Using a graduated cylinder, measure carefully approximately 25 mL 0.500 M $CuSO_4$. Record accurately the volume. Add the solution to the weighed 150- or 250-mL beaker.

3. Remove the burner under the beaker (or remove the beaker from the hot plate) and carefully pour the iron filings into the $CuSO_4$ solution a little at a time. After each addition of iron filings, wait for the reaction to subside. Record your observations.
4. To speed up the reaction, *gently* heat the beaker. The reaction should be complete when the solution reaches the boiling point. *Do not boil the liquid,* you may lose some of your product. Record your observations.
5. When the reaction is complete (no iron filings remain; look from below if you have a problem seeing whether iron filings remain), allow the solid to settle. Let the beaker cool until it is safe to touch. Add 1-2 drops of a detergent solution to prevent the solid particles from floating on the surface of the liquid. Carefully decant into a beaker (pour off the liquid leaving the solid behind). Record the appearance of the decantate (the liquid poured off). You will not be able to pour off all the liquid, but pour off as much as possible without losing any of the solid. (If you accidentally pour some of the solid off with the decantate, you can carefully pour the decantate into another beaker, wash the particles back into the beaker with the solid reaction product and try decanting again.) *Dispose of the decantate as your teacher directs.*
6. Wash the solid by adding 10 mL water (preferably distilled water). Swirl *very* gently. (Be *very* careful not to spill the solid and ruin your results!) Let the solid settle. Decant (as described in Step 5). Continue the washing procedure several times until the wash water is colorless. After the last washing decant as much water as possible so drying will not take so long.
7. Dry the solid in a drying oven. Alternatively, place the beaker where it can remain undisturbed; allow to dry overnight. (The beaker may also be placed under a heat lamp to evaporate the water.) Do not leave the solid in a drying oven overnight. If the solid is allowed to stay at an elevated temperature for a long period of time, it may react with oxygen in the air and alter the results.
8. Cool the beaker. Weigh the beaker and contents accurately and record the mass.
9. Thoroughly wash your hands before leaving the laboratory.
10. Prepare a data table like the one shown.

Data Table

1. Mass of 50-mL beaker ____________________
2. Mass of 50-mL beaker and iron filings ____________________
3. Mass of iron filings ____________________
4. Mass of 150-mL beaker ____________________
5. Mass of 150-mL beaker and $CuSO_4 \cdot 5H_2O$ ____________________
6. Mass of $CuSO_4 \cdot 5H_2O$ ____________________
7. Mass of 150-mL beaker and Cu ____________________
8. Mass of Cu ____________________

Data Table for Alternate Procedure:

1. Mass of 50-mL beaker ____________
2. Mass of 50-mL beaker and iron filings ____________
3. Mass of iron filings ____________
4. Volume 0.500 M $CuSO_4$ ____________
5. Mass of $CuSO_4$ ____________
6. Mass of 150-mL beaker ____________
7. Mass of 150-mL beaker and Cu ____________
8. Mass of Cu ____________

Part I. Data Analysis: Stoichiometry of the Reaction

1. Calculate moles of Fe used from the mass of Fe used and the molar mass of Fe.
2. Calculate moles $CuSO_4$ used. Calculate moles of $CuSO_4 \cdot 5H_2O$ from the mass of $CuSO_4 \cdot 5H_2O$ used and the molar mass of $CuSO_4 \cdot 5H_2O$. Since one mole of $CuSO_4 \cdot 5H_2O$ contains one mole of $CuSO_4$, moles of $CuSO_4 \cdot 5H_2O$ equals moles of $CuSO_4$.

   [Alternate procedure: Moles of $CuSO_4$ is calculated from the molarity and volume of $CuSO_4$: Moles = Molarity x Volume (in liters)]
3. Calculate moles Cu formed in the reaction from the mass of Cu obtained and the molar mass of Cu.
4. Tabulate the mole data for the class. Plot a graph of moles Cu *vs.* moles Fe. (Graph moles Cu on the y-axis; moles Fe on the x-axis.) Draw the best straight line through the data points.
5. Determine the slope of the line. The slope of the line gives the mole ratio (mol Cu/mol Fe). Remembering that the coefficients in the balanced equation should be integers, use this relationship between mol Cu and mol Fe to write a balanced equation for the reaction of $CuSO_4$ and Fe to form Cu and $FeSO_4$.

Part I. Implications and Applications

1. How would your experimental value for moles Cu be affected in the following situations?
   a. If 4 g $CuSO_4 \cdot 5H_2O$ were added instead of 3 g $CuSO_4 \cdot 5H_2O$? (If 30-mL 0.50 M $CuSO_4$ were added instead of 25-mL 0.50 M $CuSO_4$?)
   b. If the solid did not settle before decanting and some solid product was poured off during decanting (in Steps 5 and 6)?
   c. If the solid product was not dried thoroughly and weighed wet?
   d. If the beaker was heated strongly during drying and the solid turned black? (Strong heating can oxidize Cu to black CuO.)
2. Why was the mixture of Fe and $CuSO_4$ heated? Suggest changes that might result if the mixture is not heated.

Part II. Data Analysis: Limiting Reactant

6. From your balanced equation, what do you expect a graph of mol Cu *vs.* mol $CuSO_4$ looks like?
7. Using the class data, plot a graph of mol Cu *vs.* mol $CuSO_4$. Does the graph look like your prediction?
8. Is there a direct relationship between moles Cu formed and moles $CuSO_4$ added?
9. Using the mole ratio from the balanced equation, calculate the moles of $CuSO_4$ that would react with the total moles of Fe used. Also calculate the theoretical number of moles of Cu that would be formed from this amount of Fe.
10. Using the mole ratio from the balanced equation, calculate the moles Fe that would react with the total moles $CuSO_4$ used. Also calculate the theoretical moles Cu that would be formed from this amount of $CuSO_4$.
11. Compare the theoretical (calculated) moles Cu with the moles Cu obtained in the reaction. The reactant ($CuSO_4$ or Fe) that gives the number of moles Cu closest to the experimental value is called the *limiting reactant* since this substance limits the yield of products. Calculation of moles of product(s) from the moles of limiting reactant agrees with experimental results. What is the limiting reactant?

    The other reactant is said to be present *in excess*. The theoretical moles of this reactant is the amount that will react with the moles of limiting reactant actually used in the reaction. Some of the reactant in excess remains "left over" when the reaction is complete. What reactant is present in excess?

Part II. Implications and Applications

3. When the reaction was complete, you recorded your observations. What observation indicates that $CuSO_4$ is in excess? How would your observations differ if Fe were the excess reactant?
4. Why is the graph of mol Cu *vs.* mol $CuSO_4$ different from the graph of mol Cu *vs.* mol Fe?
5. a. Calculate the mass of $CuSO_4$ reacted.

    [Alternate procedure: Calculate the volume of 0.50 M $CuSO_4$ solution reacted.]

    b. How would your experimental results differ if you used less than the quantity of $CuSO_4$ calculated in Question 5a?
6. Calculate the mass of $FeSO_4$ that could be obtained in your reaction.
7. In most laboratory chemical reactions or in chemical manufacturing processes, there is a limiting reactant and all other reactants are present in excess. Why do you think that is so?
8. If you were in charge of a chemical plant, suggest some practical considerations that would help you to decide which reactant should be the limiting reactant.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Stoichiometry of a Reaction and Limiting Reactant

Major Chemical Concept

The purpose of this activity is to collect data that leads to establishing the mole relationship between coefficients in the balanced equation, and to introduce the concept of limiting reactant.

Level

Analysis of the data collected to determine the mole relationship between coefficients of reactants and products (*Part I*) is appropriate for all levels. Extension of the data analysis to develop the concept of limiting reactant (*Part II*) may not be appropriate for general students with very weak mathematics skills but would be appropriate for general and honors students.

Expected Student Background

For *Part I*, students should be able to make gram-to-mole conversions. They should be able to plot and read graphs and understand the concept of ratios. For *Part II*, students will also need to understand how to use the stoichiometric ratio (mole ratio) to calculate moles of product from moles of reactant.

Time

Experimental data can be collected in one class period (50 min) if a drying oven is used to dry the Cu. If the Cu is allowed to dry overnight, weighing of Cu and analysis of data can be completed on the second day. (Be sure beakers are left where they will not be disturbed and they are labeled so students can identify their beaker. Do not put stick-on labels on the beakers—this would change the mass of the beaker and invalidate the experimental results. One way to label beakers is to place a piece of paper under the beaker indicating students' names in the group and the class period.)

Safety

Read *Safety Considerations* in the *Student Version*. Although $CuSO_4$ is not considered a dangerous substance, any spills (of solution or crystals) should be cleaned up immediately. The soluble product, $FeSO_4$, is also considered reasonably safe.

Materials (For 24 students working in pairs)

Nonconsumables

12 Beakers, 150- to 250-mL
12 Beakers, 50- to 100-mL
12 Graduated cylinders, 25-mL
Balance (mg accuracy recommended)
12 Burners or hot plates
If using burner: striker or other means of lighting burner, ring stands, rings, wire gauzes
Drying oven or heat lamps for drying Cu (optional)*

Consumables

Iron filings, 6.0 g
Copper(II) sulfate, $CuSO_4 \cdot 5H_2O$, 36 g or 500 mL 0.50 M $CuSO_4$
Distilled water, 1.0 L (optional)
Detergent solution, 10 mL
*Optional: You may choose to dry the precipitate with acetone.

Advance Preparation

All equipment should be readily available for student use. If you elect to have students use 0.500 M $CuSO_4$, the solution is prepared by dissolving 62.4 g $CuSO_4 \cdot 5H_2O$ in sufficient distilled water to give 500 mL solution. You may use a graduated cylinder for measuring the quantity of the solution. Be sure all solid is dissolved before diluting to final volume. Thoroughly mix the solution and transfer to a container labeled "0.50 M $CuSO_4$." Distilled water is preferable to tap water, but the results will not be significantly affected if tap water is used.

Depending upon the number of student groups, select masses of iron filings to be used by each group. Masses should be between 0.2 and 0.5 g. This will allow for a difference in the mol Fe used and consequently the mol Cu formed. Different masses of Fe should be assigned each group to give a meaningful graph of mol Cu *vs.* mol Fe. Assign students approximate masses, *e.g.*, 0.20 g or 0.45 g. The exact mass should be determined, but it does not need to be exactly 0.200 g or 0.450 g to show the mole relationship. (The quantity of $CuSO_4$ will be in excess as long as the mass of Fe is less than 0.67 g.)

Check with your local waste disposal official about proper disposal of the $CuSO_4$ waste. $CuSO_4$ should not be flushed down the drain. Metals more active than Cu will displace Cu from $CuSO_4$. Lead can be dissolved by $CuSO_4$ that remains in plumbing traps, causing water leaks in drain lines. (The word *plumbing* comes from plumbum, lead.) Some states enforce special disposal requirements for copper salts. Collect the decantate from the reaction for proper disposal. (It should not be necessary to retain the water from the washing of the Cu, only the first decantate should require special disposal procedures.) Even if your state currently has no special disposal requirements, it is good practice to make students aware that they should not indiscriminately flush waste down the drain or throw in the garbage.

Pre-Laboratory Discussion

During the pre-laboratory, review the mole concept and the relationship between molar mass and moles of a substance. Present the Law of Conservation of Mass.

Since students will heat the reaction mixture to accelerate reaction, explain that chemical reactions occur more rapidly at elevated temperatures. As illustrations, mention more rapid souring of milk when at room temperature than under refrigeration, and the burning of gasoline in air at high temperatures whereas, it can be kept at room temperature in the presence of air without reacting.

Explain that a solution of $CuSO_4$ is used rather than the solid because reactions occur more rapidly when the reactants come into close contact. $CuSO_4$ solution contains individual Cu^{2+} and SO_4^{2-} ions. Ions are free to move about in solution and can come into closer contact with the Fe than when confined to fixed positions in the crystal.

If nomenclature has been discussed, use this opportunity to review formulas and names of reactants and products. Do not write the reaction at this point since the purpose of *Part I* of the activity is to determine the mole relationship between Fe and Cu (and by implication $CuSO_4$ and $FeSO_4$).

Review safety precautions for the use of burners if they are used for heating the reaction mixture. Generally the reaction is complete when the reaction mixture begins to boil. Caution students not to allow the solution to boil or they may lose some product, particularly if bumping occurs. (*Bumping* results from local "hot spots" in the beaker where uneven and eruptive boiling occurs. Bumping is dangerous since the hot liquid can be spewed for some distance. Beakers have been known to "jump off the ring stand" during bumping!)

Explain that each group will follow the same procedure but will use a different quantity of Fe and that all data will be combined for a class analysis of the data. It is important to accurately record the quantities of chemicals used. Problems sometimes arise when iron filings become coated with copper and thus do not fully react.

If you have not discussed solutions and plan to have students use a $CuSO_4$ solution, explain briefly the meaning of molarity (moles $CuSO_4$ per liter of solution) and why molarity is a convenient way of expressing the solution's composition (easy to determine moles $CuSO_4$ by measuring the solution volume). You may prefer to lead the entire class in a calculation of moles $CuSO_4$ used in the reaction. A simple way to explain this would be to say: "If 1 L or 1000 mL contains 0.500 mol $CuSO_4$, how much $CuSO_4$ would 100 mL contain?

$$100 \text{ mL} = \frac{1}{10} \text{ of } 1000 \text{ mL, thus } \frac{1}{10} \text{ of } 0.500 \text{ mol} = 0.0500 \text{ mol } CuSO_4.$$

If 100 mL contains 0.0500 mol $CuSO_4$, how much $CuSO_4$ would 25 mL contain?

$$25 \text{ mL} = \frac{1}{4} \text{ of } 100 \text{ mL, thus } \frac{1}{4} \text{ of } 0.0500 \text{ mol} = 0.0125 \text{ mol } CuSO_4."$$

You might tell students how many moles (or grams) of $CuSO_4$ are in 25 mL of the $CuSO_4$ solution (0.0125 mol or 2.00 g $CuSO_4$). (*NOTE: If your students weigh their own* $CuSO_4 \cdot 5H_2O$ *and dissolve it, about 3 g is required because the inclusion of the water of hydration in* $CuSO_4 \cdot 5H_2O$ *increases the molar mass from 159.7 g/mol to 249.7 g/mol.*)

If a drying oven is used, the sample should not be dried overnight as Cu oxidation may occur at the elevated temperature.

Teacher-Student Interaction

Move through the groups insuring proper safety procedures. Try not to answer many questions. Answer a question with a question when possible to facilitate developing thinking skills. Possible questions are:

1. What does the blue decantate indicate? *[Since* $CuSO_4$ *is used in excess, when the reaction is complete there will be some* $CuSO_4$ *remaining. Be sure students note the color of the decantate. This is one of the follow-up questions.]*
2. Why is heating continued after iron filings are added? *[To speed up the reaction.]*
3. Why is a solution of $CuSO_4$ used rather than just mixing crystals with iron filings? *[This is discussed in the* Pre-Laboratory Discussion *section.]*
4. Why does the activity call for iron filings instead of iron nails or other large pieces of iron? *[Because reaction is more rapid for finely divided solid.]*
5. Why is the product washed? *[To remove impurities,* $FeSO_4$ *and excess* $CuSO_4$*, from the Cu.]*

If students weigh $CuSO_4 \cdot 5H_2O$ and dissolve it, they will find that it does not dissolve readily.

6. How can you speed up the dissolving process for $CuSO_4 \cdot 5H_2O$? *[Stirring, grinding the crystals, heating the solution. Stirring and heating are methods used in this activity.]*
7. Why might grinding the crystals in a mortar and pestle not be a good idea? *[Unless the crystals are ground before weighing, the mass weighed and the mass dissolved might not be the same since some crystals might be lost (left on mortar and pestle) during crushing and not transferred to the solution.]*

Anticipated Student Results

Sample Data Table

1.	Mass of 50-mL beaker	26.292 g
2.	Mass of 50-mL beaker and iron filings	26.603 g
3.	Mass of iron filings	0.311 g
4.	Mass of 150-mL beaker	98.325 g
5.	Mass of 150-mL beaker and $CuSO_4 \cdot 5H_2O$	101.367 g
6.	Mass of $CuSO_4 \cdot 5H_2O$	3.042 g
7.	Mass of 150-mL beaker and product (Cu)	98.673 g
8.	Mass of dry product (Cu)	0.348 g

Part I. Data Analysis: Stoichiometry of the Reaction

1. mol Fe: $0.311 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 5.57 \times 10^{-3} \text{ mol Fe}$

2. mol $CuSO_4$: $3.042 \text{ g } CuSO_4 \cdot 5H_2O \times \frac{1 \text{ mol } CuSO_4 \cdot 5H_2O}{249.70 \text{ g } CuSO_4 \cdot 5H_2O} =$

1.218×10^{-2} mol $CuSO_4 \cdot 5H_2O$

1.218×10^{-2} mol $CuSO_4 \cdot 5H_2O$ = 1.218×10^{-2} mol $CuSO_4$

If the alternate procedure is used, mol $CuSO_4$:

$$0.0252 \text{ L solution} \times \frac{0.500 \text{ mol } CuSO_4}{1 \text{ L solution}} = 1.26 \times 10^{-2} \text{ mol } CuSO_4$$

3. $0.348 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.54 \text{ g Cu}} = 5.48 \times 10^{-3} \text{ mol Cu}$
4. (Graph)
5. Slope should be approximately 1, this is, 1 mol Cu : 1 mol Fe.

Balanced equation: $Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + FeSO_4(aq)$

Part II. Data Analysis: Limiting Reactant

6. Graph predicted is a straight line with slope 1 (1 mol Cu/1 mol $CuSO_4$).
7. Graph does not look like prediction. Line is almost vertical.
8. There is no direct relationship between moles Cu formed and moles $CuSO_4$. The moles Cu vary even though moles $CuSO_4$ are essentially constant for all groups.
9. mol $CuSO_4$ and mol Cu ("theoretical if all Fe is used"):

$$5.57 \times 10^{-3} \text{ mol Fe} \times \frac{1 \text{ mol } CuSO_4}{1 \text{ mol Fe}} = 5.57 \times 10^{-3} \text{ mol } CuSO_4$$

$$5.57 \times 10^{-3} \text{ mol Fe} \times \frac{1 \text{ mol Cu}}{1 \text{ mol Fe}} = 5.57 \times 10^{-3} \text{ mol Cu}$$

10. mol Fe and mol Cu ("theoretical if all $CuSO_4$ is used"):

$$1.218 \times 10^{-2} \text{ mol } CuSO_4 \times \frac{1 \text{ mol Fe}}{1 \text{ mol } CuSO_4} = 1.218 \times 10^{-2} \text{ mol Fe}$$

$$1.218 \times 10^{-2} \text{ mol } CuSO_4 \times \frac{1 \text{ mol Cu}}{1 \text{ mol } CuSO_4} = 1.218 \times 10^{-2} \text{ mol Cu}$$

11. Fe is the limiting reactant. $CuSO_4$ is present in excess.

A table such as the following may be useful in comparing the experimental and theoretical results.

Comparison of Theoretical and Experimental Yields of Cu

	Experimental	Theoretical (If all Fe is used)	Theoretical (If all $CuSO_4$ is used)
mol Fe	5.57×10^{-3} mol	5.57×10^{-3} mol	1.218×10^{-2} mol
mol $CuSO_4$	1.218×10^{-2} mol	5.57×10^{-3} mol	1.218×10^{-2} mol
mol Cu	5.48×10^{-3} mol	5.57×10^{-3} mol	1.218×10^{-2} mol

Answers to Implications and Applications

Part I. Stoichiometry of the Reaction

1. a. No effect. $CuSO_4$ is in excess with 3 g. Adding more would not affect the yield of Cu.

   b. Experimental mol Cu would be less because Cu would be poured off with decantate and not recovered and weighed.

   c. Experimental mol Cu would be more because water added to the mass of Cu. The total mass of Cu and water was used to calculate mol Cu.

   d. Experimental mol Cu would be more because Cu was oxidized to CuO, which would have a greater mass than Cu. Since the total mass (Cu and combined O) was used to calculate mol Cu, the apparent mol Cu would be greater.

2. The mixture was heated to speed up the reaction. If the reaction mixture were not heated, the reaction would take longer. If the reaction did not go to completion before decanting, some Fe might remain unreacted and be weighed as Cu, causing an error in mol Cu.

Part II. Limiting Reactant

3. The blue color of the solution indicates $CuSO_4$ is present after the reaction is complete. If Fe were the reactant in excess, the decantate would not be blue. (Also Fe would be mixed with Cu and the mass of the remaining solid would not be the mass of Cu. Do not emphasize this point, however, unless you want to discuss how the activity could be modified to determine the mass of Cu. One possibility is to dissolve the excess Fe in HCl to recover Cu, which is insoluble in HCl.)

4. Since $CuSO_4$ is in excess, there is not a mole relationship between mol $CuSO_4$ added and mol Cu formed. The mole relationship exists only for the limiting reactant (Fe).

5. a. Mass $CuSO_4 \cdot 5H_2O$ (volume 0.500 M $CuSO_4$):
   1 mol $CuSO_4$ = 1 mol $CuSO_4 \cdot 5H_2O$

$$5.57 \times 10^{-3} \text{ mol } CuSO_4 \cdot 5H_2O \times \frac{249.7 \text{ g } CuSO_4 \cdot 5H_2O}{1 \text{ mol } CuSO_4 \cdot 5H_2O} = 1.39 \text{ g } CuSO_4 \cdot 5H_2O$$

$$[5.57 \times 10^{-3} \text{ mol } CuSO_4 \times \frac{1 \text{ L solution}}{0.500 \text{ mol } CuSO_4} = 0.0111 \text{ L} = 11.1 \text{ mL solution}]$$

   b. $CuSO_4$ would be the limiting reactant; the experimental mass (and amount in moles) Cu recovered would be less.

6. Mass $FeSO_4$:

$$5.57 \times 10^{-3} \text{ mol Fe} \times \frac{1 \text{ mol } FeSO_4}{1 \text{ mol Fe}} \times \frac{152.9 \text{ g } FeSO_4}{1 \text{ mol } FeSO_4} = 0.852 \text{ g } FeSO_4$$

7. Many students recognize that it is difficult to measure exactly equivalent mole amounts of all reactants. (Another reason is that adding an excess of one reactant insures that all of the other reactant is used in the reaction.)

8. Practical considerations include: Which reactant is easier to measure? Which reactant is less expensive?

Post-Laboratory Discussion

Reinforce the significance of moles in chemical reactions. Be sure students understand that coefficients in an equation relate to moles (not grams) of reactants and products. Discuss the concept of limiting reactant. Analogies are particularly helpful (see *Metaphors and Analogies*). Review student responses to the questions in *Implications and Applications.* Item 1 in *Possible Extensions* could be used during post-laboratory discussion.

Possible Extensions

1. An alternative treatment of data is to have students calculate the ratio mol Cu/mol Fe and plot the number of students having a given mole ratio (on the y-axis) *vs.* the mole ratio (on the x-axis). Mole ratios should be rounded to ± 0.1; *i.e.*, plot ratios of 0.7, 0.8, 0.9, 1.0, 1.1, *etc.* It should be clear that most data (the middle two-thirds) fall very close to 1.0. The range of the middle two-thirds of the mol Cu/mol Fe data can be noted and compared with uncertainty in the measurements.
2. Repeat the activity with other metals such as Al, Mg, Zn. Mole ratios can be determined and equations written. (Adjustments in the quantity of metal may have to be made to insure that $CuSO_4$ is present in excess.) In Activity S5 "What Is the Copper Formula?" in *Laboratory Assessment Builds Success* the percent copper in $CuCl_2$ is determined by reaction of Al with $CuCl_2$. The activity could be adapted to illustrate stoichiometry if the data is treated as in this activity.
3. If Mg or Zn is used and students are told that reaction stoichiometry is the same as for the reaction of $CuSO_4$ with Fe, the molar mass of the metal can be determined from data collected (mass of metal reacted and mass of Cu formed).
4. *Activity 1: Limiting Reagent for a Chemical Reaction in Solution* in the *Solutions* module illustrates solution stoichiometry.

Assessing Laboratory Learning

1. **Laboratory Practical.** Give students sample data (mass or mol Fe and $CuSO_4$) and ask them to determine which is the limiting reagent and the theoretical yield of Cu. Give them data for mol of unknown metal (M) and mol $CuSO_4$, and ask them to plot the data to determine the reaction stoichiometry. (Alternately, a graph of mol unknown metal (M) *vs.* mol Fe could be given and students asked to determine the stoichiometry.)
2. **Laboratory Report.** Although the data for the entire class are used for the graph to determine the reaction stoichiometry, it is worthwhile to have students plot their own graph as part of the activity report. Explain that graphing is important in analysis of chemical data and that they should become familiar with techniques for plotting and reading graphs. Also important is the idea that use can be made of the same data for different purposes: to determine the stoichiometry of the reaction and to demonstrate which reactant is the limiting reactant. Because the data analysis is somewhat involved, you may divide the laboratory report into two parts or use an abbreviated format. An abbreviated format might consist of tables, graphs and brief discussion of the meaning of the graphs and tables with calculations in an appendix.
3. **Written Examination.** Some possible items:
   a. An unknown metal reacts with an excess of $CuSO_4$ to form Cu. The following data are determined for reaction of the unknown metal with $CuSO_4$:

mol metal(M) reacted	mol Cu formed
2.80×10^{-3}	4.20×10^{-3}
3.70×10^{-3}	5.55×10^{-3}
4.11×10^{-3}	6.17×10^{-3}
5.52×10^{-3}	8.28×10^{-3}

   Plot mol Cu *vs.* mol M and find the coefficients of M (the unknown metal) and Cu in the balanced equation. *[3 mol Cu : 2 mol M]*

   b. When 0.859 g Fe is reacted with 3.47 g $CuSO_4$, what is the theoretical yield of Cu? *[0.979 g Cu—Fe is the limiting reactant.]*

References

Cotton, F. A., Darlington, C. L., and Lynch, L. D. (1980). *Chemistry: An investigative approach.* Boston, MA: Houghton Mifflin.

Gardner, M., Smith, P., and Heikkinen, H. (1990). *Laboratory assessment builds success.* Berkeley, CA: Lawrence Hall of Science, Institute for Chemical Education.

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Percent Yield of a Chemical Reaction

Introduction

The yield for a chemical reaction, the relationship between that yield, and the balanced equation representing the reaction are very important to chemists. Not only does a laboratory chemistry investigation require an understanding of such concepts, but the chemical industry also depends on percent yield principle in the manufacture of chemical products.

The reaction in this activity uses sodium bicarbonate and hydrochloric acid. The equation for this reaction shows the mole relationships for the reaction. The yield of NaCl will be determined and compared with the theoretical yield.

$$HCl(aq) + NaHCO_3(s) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Purpose

To illustrate the relationship between the balanced equation and the amounts of substances involved in a reaction either as reactants or products. To determine the percent yield of a reaction.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Exercise caution when adding the HCl solution to sodium bicarbonate to avoid splashing during reaction. When drying the product, heat slowly and cautiously to avoid spattering. HCl can cause severe burns and can irritate the respiratory tract.
3. If burners are used to evaporate the solution, exercise caution in lighting the burner. Watch out for loose hair. Hair ignites easily especially if hair spray is used.
4. Always exercise caution when handling glassware, especially if the glassware is hot. Hot and cold glass look the same!

Procedure

1. Weigh a clean dry 250-mL beaker.
2. Add approximately 2.5 g $NaHCO_3$ to the beaker.
3. Accurately weigh the beaker and $NaHCO_3$.
4. Slowly add the HCl solution (about 10-15 mL) and observe. You may be told to use a dropper to add the acid or just pour very slowly.
5. When the reaction has apparently stopped, add a few more milliliters of the HCl solution to be sure all $NaHCO_3$ has reacted.
6. Evaporate the solution to recover the product, NaCl. Your teacher will tell you how to evaporate the resulting solution to obtain dry NaCl. You may use a drying oven or a burner. If a burner is used, heat slowly and carefully to avoid spattering. Heat in a hood. Be careful not to breathe fumes.
7. Thoroughly wash your hands before leaving the laboratory.

Observations

Make observations about substance color, nature of the reaction, *etc.*, as you perform the laboratory.

Data Analysis

1. Prepare a data chart as shown:

   Mass of empty beaker ______________

   Mass of beaker and $NaHCO_3$ ______________

   Mass of $NaHCO_3$ ______________

   Mass of beaker and NaCl after evaporating ______________

   Mass of NaCl recovered ______________

2. From grams $NaHCO_3$ used, calculate mol $NaHCO_3$ used and mass NaCl formed (theoretical yield) using the balanced equation.
3. Using the mass NaCl recovered, calculate mol NaCl recovered.
4. Calculate the percent yield for the reaction from the actual yield (g NaCl recovered) and the theoretical yield (g NaCl calculated in Question 2).

$$\%\ \text{Yield} = \frac{\text{Actual yield recovered}}{\text{Theoretical yield}} \times 100$$

5. Share class data. Suggest reasons why the percent yield does not equal 100%. (It may be greater than 100%!)

Implications and Applications

1. Why can the percent yield *theoretically* not be greater than 100%? Use the Law of Conservation of Mass to support your answer.
2. Why is it necessary to add sufficient HCl to ensure the complete reaction of $NaHCO_3$?
3. What could happen to the percent yield in each of the following situations:
   a. The mass of $NaHCO_3$ used is recorded as 2.345 g; it is actually 2.545 g?
   b. The mass of NaCl recovered is recorded as 1.786 g; it is actually 1.886 g?
   c. Some distilled water is accidentally added to the beaker containing NaCl before the solution is evaporated?
4. What effect would using a more dilute HCl solution have on the percent yield?

*5. Suppose $NaHCO_3$ used in this activity is contaminated with NaCl. How would this affect the percent yield? (See comments in *Teacher Notes.*)

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Percent Yield of a Chemical Reaction

Major Chemical Concept

Data collection to support the stoichiometric relationships based on the balanced equation, and the introduction of the percent yield concept are the main goals of this activity. The emphasis is on the relationship between the balanced equation and the actual amounts of substances used and produced in this chemical reaction. The concept of limiting reactant while not stressed in this activity is implied in the part where students are instructed to add additional HCl after the reaction had apparently stopped. One of the student questions relates to this. It is *not* too early to introduce the idea of excess and limiting reactants. Comparison of class data for this activity shows that percent yields are independent of the amounts of reactants.

One major weakness of this activity is all mole ratios are 1:1, which may tend to lead to a careless assumption that moles of reactant are always equal to moles of product. This laboratory should be followed by a thorough discussion of ratios other than 1:1, as well as those that are 1:1.

Level

All levels of introductory chemistry should be able to complete this activity, with varying amounts of teacher input and guidance. (Question 5 in *Implications and Applications* may be too difficult for general-level students.)

Expected Student Background

Students should be able to make mole conversions (grams to moles) and to compute percentage. Ratio and proportion concepts will be needed to connect actual molar amounts used in the laboratory and mole ratios according to the balanced equation.

Time

This activity can be completed easily in less than 50 min if the proper pre-laboratory is provided and a drying oven is used to evaporate the solution containing the NaCl. If burners are used to evaporate the solutions, finish the reaction one day and complete the evaporation the next. If a drying oven is used, weighing the next day takes very little time. This allows ample class time to collect class data and make calculations.

Safety

Read the *Safety Considerations* in the *Student Version*.

Materials (For 24 students working in pairs)

Nonconsumables

- 12 Beakers, 250-mL
- Balance (mg accuracy recommended)
- 12 Burners or drying oven
- If using burner: striker or other means of lighting burner, ring stands, rings, wire gauzes
- If drying oven is used: aluminum foil

Consumables

- Sodium bicarbonate, $NaHCO_3$, 24-36 g
- 3 M Hydrochloric acid, HCl, 250 mL (62.5 mL conc. HCl diluted to 250 mL)

Advance Preparation

Prepare 3 M HCl and place in dropper bottles for each laboratory team.

Check all equipment ahead to be sure all is in good working condition. Having equipment in plain view at laboratory stations will save much time! *WARNING: If a drying oven is used, cover the beaker with aluminum foil.* HCl vapors will react with the walls of many drying ovens causing rapid oxidation (rusting). Any aluminum chloride formed by the reaction of HCl with aluminum foil clings to the aluminum foil and will not affect experimental results.

Decide whether to assign Question 5 in *Implications and Applications* or use it in the post-laboratory discussion.

Pre-Laboratory Discussion

The purpose of this activity is to study stoichiometry, not to predict reaction products. Review stoichiometric relationships and the calculations required in this activity. During the pre-laboratory, consider demonstrating the addition of HCl solution to $NaHCO_3$. Have students feel the beaker before and after adding HCl to $NaHCO_3$. It should become noticeably cooler. Proper laboratory skills, such as lighting burners, should be stressed. Thorough safety instructions must be given. If you choose to omit Question 5 in *Implications and Applications*, inform the students.

Possible pre-laboratory discussion questions are:

1. Why is additional HCl solution added? *[To insure the reaction is complete.]*
2. Which product might be recovered easily for purposes of measuring the amount? *[Consider the physical states of the products. It should be obvious that carbon dioxide gas would be difficult but not impossible to recover. It is not feasible to try to determine how much water is made since the amount is so small compared to the amount of water in the HCl solution. NaCl would be the easiest product to recover.]*
3. What data should be collected? *[See data chart in* Student Version.*]*
4. How could you use collected data to determine the stoichiometry of the reaction? *[See* Activity 1: Stoichiometry of a Reaction and Limiting Reactant.*]*
5. How could you use the data to determine percent yield? *[This is described in Questions 2-4 in* Data Analysis *of* Student Version. *Be sure to include in your discussion the need to collect class data for comparison of results.]*

Teacher-Student Interaction

Move among groups to insure proper safety procedures. Try not to answer many questions. Answer a question with a question when possible to facilitate developing thinking skills. Possible questions are:

1. Did you notice a temperature change in the beaker during the reaction? *[Reaction is endothermic—beaker gets cool. Although most neutralization reactions are exothermic, this reaction is endothermic. Students may have observed that the glass gets cold when Alka-Seltzer (principally citric acid and* $NaHCO_3$*) is dissolved in water. You might wish to point out that this activity is essentially the same reaction (Acid + Bicarbonate).]*
2. What evidence do you observe for a chemical reaction? *[Fizzing, beaker gets cool.]*
3. What causes the fizzing? *[*$CO_2(g)$*]*

This laboratory can be very quickly completed leaving time for the discussion about observations.

Anticipated Student Results

An actual data chart from a student's laboratory report:

Mass of empty beaker	93.650 g
Mass of beaker and $NaHCO_3$	95.151 g
Mass of $NaHCO_3$	1.501 g
Mass of beaker and NaCl after evaporating	94.691 g
Mass of NaCl	1.041 g

Answers to Data Analysis

1. An actual data chart from a student's laboratory report is shown in *Anticipated Student Results*.
2. $$1.501\text{ g NaHCO}_3 \times \frac{1\text{ mol NaHCO}_3}{84.01\text{ g NaHCO}_3} = 1.787 \times 10^{-2}\text{ mol NaHCO}_3$$
   $$1.501\text{ g NaHCO}_3 \times \frac{1\text{ mol NaHCO}_3}{84.01\text{ g NaHCO}_3} \times \frac{1\text{ mol NaCl}}{1\text{ mol NaHCO}_3} \times \frac{58.44\text{ g NaCl}}{1\text{ mol NaCl}} = 1.044\text{ g NaCl}$$
3. $$1.041\text{ g NaCl} \times \frac{1\text{ mol NaCl}}{58.44\text{ g NaCl}} = 1.781 \times 10^{-2}\text{ mol NaCl}$$
4. $$\frac{1.041\text{ g NaCl}}{1.044\text{ g NaCl}} \times 100\% = 99.7\%\text{ yield}$$
5. Spattering during the reaction or during evaporation would give low results. Water (sample is not dry) or other impurities could cause high results. (Further drying should improve results if the sample is wet.) Incomplete reaction of $NaHCO_3$ will also give high results. Incorrect masses could result in either high or low results.

Answers to Implications and Applications

1. The percent yield theoretically should not be greater than 100% because the Law of Conservation of Mass states that matter is not created or destroyed in ordinary chemical reactions. The concept of a balanced equation is based on the Law of Conservation of Mass.
2. Theoretical yield is based on the mass of $NaHCO_3$ added.
3. a. Since mass of $NaHCO_3$ is too small, the theoretical yield of NaCl will be too small. The calculated percent yield will be too high (probably greater than 100%).
   b. The actual yield of NaCl is too small. The calculated percent yield will be too low.
   c. There would be no effect on the percent yield since the water is evaporated.
4. None. It would simply require more HCl solution to complete the reaction.
5. If $NaHCO_3$ is contaminated with NaCl, the assumed mass of $NaHCO_3$ would be too large, making the theoretical yield of NaCl too large. Thus the calculated percent yield will be too low. In discussing this with students, select masses for $NaHCO_3$ and NaCl (*e.g.*, 1.00 g NaCl and 1.00 g $NaHCO_3$) and calculate "theoretical yield" (assuming 2.00 g $NaHCO_3$) and "actual yield" (assuming 1.00 g NaCl and 1.00 g $NaHCO_3$).

Post-Laboratory Discussion

Post-laboratory discussion should start by collecting class data. Collect the data for moles $NaHCO_3$ used, the moles NaCl produced, and the percent yield for each group. The class should be able to determine that percent yield is independent of how much reactant is used. Use moles NaCl collected compared to moles NaCl predicted as well as grams collected compared to grams predicted to calculate the percent yield. This is another way of reinforcing the concept that mole is an amount. This activity is fairly consistent in giving very close to 100% yields. Discussing how the mole ratio from the balanced equation is used in this laboratory is essential. Ask such questions as "Does the amount of $NaHCO_3$ used affect the amount of NaCl produced?" and "Why does the mass of NaCl characteristically seem to be less than the mass of $NaHCO_3$ used if the mole ratio is 1:1?" This leads to the discussion of different molar masses for compounds.

Propose the idea that the students are chemists in a chemical manufacturing plant in charge of deciding how much of each reactant they need to have on site for the production of certain chemicals. How would percent yield for their product enter into the decision about amounts of reactants to order?

Extensions

1. An important part of chemical education is the continued integration of all disciplines in the learning process. After collecting the data for moles $NaHCO_3$ and moles NaCl involved in this reaction, have students plot a graph of mol $NaHCO_3$ *vs.* mol NaCl. The slope of the line is close to one as expected from the balanced equation. This discussion should lead to the question, if the ratio is not 1:1 for a particular comparison being made, what happens to the slope? Can the slope of such a line be used to determine mole ratios of reactions?
2. If you have discussed gases, you might ask students to suggest ways to determine how much CO_2 is generated in the reaction. The most frequent answer is to collect the gas by water displacement. Discuss why this would not be practical for CO_2. *[CO_2 dissolves in water as evidenced by its use to "carbonate" beverages.]* Students could research methods for determining CO_2.
3. Students need to be aware that data is not valid unless statistically evaluated. Simple statistical analysis could involve calculating the average and standard deviation of class data. Students do not need to understand the theory behind the formulas used to derive standard deviation. Choose one of the available computer software packages for the calculations. Connect the average and standard deviation to the normal curve to give students an appreciation for the ranges that exist for measured quantities. (The mathematics and science departments could team-teach here.)

Assessing Laboratory Learning

1. **Laboratory practical.** Provide students with sample data from a different laboratory activity and have them make the appropriate calculations to determine percent yield. Try to use data that requires a different approach to the same concepts.

   $$MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$$

   500 mL of 0.429 M solution $MgSO_4$ was treated with 0.500 M NaOH.

a. How much NaOH was necessary to precipitate the $Mg(OH)_2$ completely? *[858 mL]*

b. Assuming 100% yield, how many g of $Mg(OH)_2$ can be expected to form? *[12.5 g]*

c. The actual yield for this reaction is 10.4 g. What is the percent yield? *[83.2%]*

d. Can you think of a reason why the percent yield is so low? *[$Mg(OH)_2$, although only slightly soluble in water, has a fairly large solubility product constant (see* Solubility and Precipitation *module). If a filtration method is used to separate the precipitated $Mg(OH)_2$ from the supernatant liquid, an appreciable amount of $Mg(OH)_2$ will remain in solution, contributing to the low yield.]*

2. **Laboratory report.** One variation on the individual laboratory report is a group laboratory report. With the latter, the group consists of perhaps 4 or 5 pairs of students who worked separately. After completing the laboratory, all pairs combine data and analyze the outcome. Requirements can be more stringent. The laboratory report must be typed, double spaced, *etc.* The grading should be done on a very competitive basis. It is better to have another competent and trusted teacher evaluate the reports on the basis of information presentation. This makes the activity more realistic because in real-world situations groups of individuals combine efforts for a final product, often with outside evaluation.

3. **Examination questions.** Some possible items:

a. Three samples of a given metallic ore are weighed. The masses are 12.45 g, 34.65 g, and 78.90 g, respectively. Analysis of the ores reveals that the samples contain 7.22 g, 21.14 g, and 44.97 g of copper, respectively. Based on these data, if a metallurgist wished to reclaim the copper from 125 kg of this ore, and knew that the process required 4 mol of O_2 gas for every 2 mol of copper to be recovered, how many mol O_2 gas would be needed? *[The % Cu in the ore is 58.7% (average of 3 analyses that give 57.0%, 61.01% and 57.00%, respectively), 125 kg ore would contain 73.4 kg Cu (1.15×10^3 mol Cu) and would require 2.30×10^3 mol O_2.]*

b. A reaction has a percent yield of 36.7%. If the reaction involves the production of CO_2 gas at this percent yield, how many moles of Reactant A would be needed to give 88.0 g of CO_2 if the mole ratio is 3 mol A : 1 mol CO_2?

$$88.0\ g\ CO_2 \times \frac{1\ mol\ CO_2}{44.0\ g\ CO_2} \times \frac{3\ mol\ A}{1\ mol\ CO_2} = 6.00\ mol\ A$$

$$36.7\%\ X = 6.00\ mol\ A$$

$$X = \frac{6.00\ mol}{0.367} = 16.3\ mol\ A$$

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Balancing Chemical Equations

Purpose

To illustrate the relationship of the Law of Conservation of Mass to balanced chemical equations.

Materials

- Gumdrops
- Toothpicks

Safety

No special precautions are needed. Students should not eat the candy, particularly if this is done in the laboratory.

Directions

Using gum drops (or jelly beans) for atoms and toothpicks to hold atoms together, prepare models of reactant compounds ("reactant molecules"). Different colors of gum drops should be used to represent different elements. Several "molecules" of each reactant should be made. The atoms are then rearranged to form the product "molecules." All gum drops from any "molecule" that reacts must be used. None may be lost (eaten?!) and "extra atoms" not in the reactant "molecules" may not be used to make products. When the "product molecules" have been made, the reaction is balanced. The balanced equation is written for the "reaction."

The purpose is to emphasize the Law of Conservation of Matter in balancing equations. The arrangement of atoms in the molecules is unimportant. The purpose is to illustrate the Law of Conservation of Mass in balancing equations, not molecular geometry. Students should be given several *unbalanced* reactions for which to make models and write a balanced equation. Examples of appropriate reactions (unbalanced) are:

$$\begin{array}{rcl}
H_2 + O_2 & \rightarrow & H_2O \\
NH_3 & \rightarrow & N_2 + H_2 \\
H_2S + SO_2 & \rightarrow & H_2O + S \\
P + Cl_2 & \rightarrow & PCl_5 \text{ (or } PCl_3\text{)} \\
CH_4 + O_2 & \rightarrow & CO_2 + H_2O
\end{array}$$

(or combustion reaction for other hydrocarbons, such as C_2H_6, C_2H_4, C_3H_8)

Some more difficult reactions to balance:

$$\begin{array}{rcl}
NO_2 + O_2 & \rightarrow & N_2O_5 \\
NH_3 + O_2 & \rightarrow & NO + H_2O \\
PBr_3 + H_2O & \rightarrow & H_3PO_3 + HBr \\
PCl_5 + H_2O & \rightarrow & H_3PO_4 + HCl \\
SCl_4 + H_2O & \rightarrow & H_2SO_3 + HCl
\end{array}$$

Demonstration 2: Electrolysis of Water

Purpose

To illustrate reaction stoichiometry involving gaseous products.

Materials

- Electrolysis apparatus (The *Electrochemistry* module has information on electrolysis and construction of electrolysis apparatus.)
- Water
- 6 M Sulfuric acid, H_2SO_4, 10 mL
- Wooden splint

Safety

Do not allow the reaction to proceed long enough to collect large quantities of hydrogen gas. Hydrogen gas can be explosive if ignited.

Directions

Electrolysis of water will yield two volumes of hydrogen for every volume of oxygen. Since water will not conduct an electric current, a small amount of an electrolyte that will not be oxidized or reduced as easily as water (usually sulfuric acid) is added to the water to provide ions to carry the charge. The identity of the two gases can be determined by showing that a glowing splint will burn when thrust into the oxygen, but not the hydrogen.

Demonstration 3: Limiting Reactant

Purpose

To illustrate the limiting reactant concept.

Materials

Household vinegar (5% acetic acid, $HC_2H_3O_2$), 100 mL
Sodium bicarbonate, $NaHCO_3$ (baking soda), 5 g
Funnel
6 Party balloons
6 Test-tubes, 18- x 150-mm

Safety

No special safety precautions are needed

Directions

1. Add 10.0 mL of household vinegar to each of six 18- x 150-mm test-tubes. Place these test-tubes in a test-tube rack.
2. Obtain six small round party balloons, the same size and preferably the same color.
3. Weigh one of the balloons. Using a funnel add 0.18 g of household baking soda. Be sure that the baking soda goes to the bottom of the balloon and none remains near the opening.
4. Repeat Step 3 for the remaining five balloons containing 0.35 g, 0.52 g, 0.70 g, 1.00 g and 1.70 g, respectively. (The balloons may be prepared ahead of time for demonstration the following day.)
5. Attach each of the six balloons to the test-tubes prepared in Step 1, taking care not to mix the contents of the balloons and test-tubes (see Figure 1).
6. After the balloons are securely attached to the test-tubes, lift the balloons one at a time and allow mixing of the contents of the balloon and the test-tubes. Enlist student assistants for this part of the demonstration, especially to secure the balloon to the test-tubes during inflation.

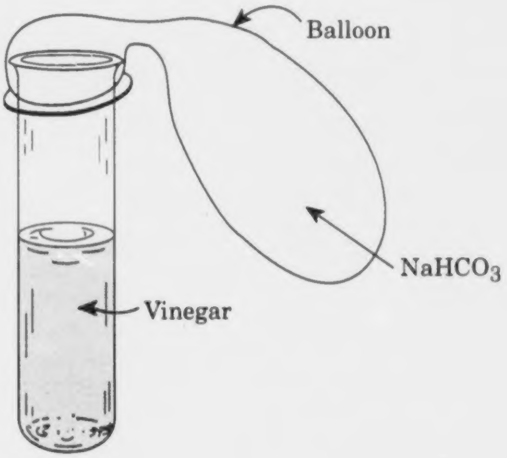


Figure 1. Diagram for limiting reactant demonstration.

Observations: Balloons should inflate with the gaseous product of this reaction. Balloons attached to Tubes 1 though 4 should gradually increase in volume. Balloons attached to Tubes 4 through 6 should be approximately the same size, even though more baking soda has been added to these balloons.

1. Which reactant was limiting in test-tube . . .

   1? 2? 3?

   4? 5? 6?

   [Tubes 1-3: $NaHCO_3$; Tubes 5-6 $HC_2H_3O_2$; Tube 4 has approximately stoichiometric proportions of the reactants.]

   Reaction occurring is: $2\,H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

Calculations:

2. Household vinegar is 5% acetic acid, $HC_2H_3O_2$ (MM = 60 g/mol). If the density of household vinegar is approximately 1.00 g/mL, how many moles of acetic acid are contained in 10 mL of household vinegar? *[8.3×10^{-3} mol]*

3. Find the number of moles of sodium bicarbonate, $NaHCO_3$ (MM = 84 g/mol; baking soda) contained in each of the balloons:

Balloon 1: 0.18 g ____ *[2.1×10^{-3} mol]*
Balloon 2: 0.35 g ____ *[4.2×10^{-3} mol]*
Balloon 3: 0.52 g ____ *[6.2×10^{-3} mol]*
Balloon 4: 0.70 g ____ *[8.3×10^{-3} mol]*
Balloon 5: 1.00 g ____ *[1.2×10^{-2} mol]*
Balloon 6: 1.70 g ____ *[2.0×10^{-2} mol]*

4. Consider answers to calculation Questions 2 and 3 and your observations about limiting reactant in Question 1. Do they agree? Explain the relationship of the calculations to your observations.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. How is the Law of Conservation of Mass related to a balanced equation? *[In a balanced equation the mass of reactants equals the mass of products.]*
2. How are moles related to reaction stoichiometry? *[A balanced equation relates moles of products to moles of reactants. Since moles is related to mass of a substance, masses of reactants and products can be related.]*
3. What is the relationship between limiting reactant and theoretical yield? *[Since the limiting reactant is present in smallest stoichiometric amount, it determines theoretical yield.]*
4. How does percent yield support the idea that actual yield cannot be more than theoretical yield? (How does this relate to the Law of Conservation of Mass?) *[Because mass is conserved in all chemical reactions, percent yield can not be greater than 100%. This implies that the theoretical yield is the maximum yield possible; thus, actual yield must be no more than the yield calculated from the stoichiometric relationship. (If actual yield appears greater, it is because of experimental error or because impurities are present.)]*
5. Why does the calculation of the theoretical yield imply that the reaction goes to completion? *[If the reaction does not go to completion, reactants will remain, and the actual yield will be less than the calculated yield.]*

Counterintuitive Examples and Discrepant Events

The puzzling nature of chemical reactions is illustrated by bubbling carbon dioxide into a calcium hydroxide solution. Initially a precipitate ($CaCO_3$) will form. Upon continued addition of carbon dioxide, the precipitate disappears due to formation of calcium bicarbonate, which is more soluble. This is a stoichiometry effect:

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

$$CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$$

For further discussion see *Acids and Bases* module.

Analogies and Metaphors

Limiting Reactant

1. **Tricycle**: How many tricycles can be made from a given number of wheels, handle bars, frames, seats, *etc.*? What will be left? ***Example***: How many tricycles can you make from 10 wheels, 5 handle bars, 4 frames and 6 seats? *[3]*

2. **Making S'mores**: Purchase a package of six Hershey's™ chocolate bars, a box of graham crackers, and a bag of marshmallows.

   Recipe for making S'mores:

   3 chocolate squares
   1 graham cracker (broken into two parts)
   2 toasted marshmallows

   The package of chocolate bars contains 6 bars, each bar has twelve chocolate squares. Each box of graham crackers contains three packages of ten crackers (may vary). Each bag of large marshmallows contains 80 marshmallows (may vary).

   Ask students the following questions:

   a. What is the maximum number of S'mores that can be made? *[24]*

   b. Which ingredient is limiting? *[Chocolate]*

   c. How many each of the other ingredients will be left? *[6 graham crackers, and 32 marshmallows]*

   d. If we wanted to make more S'mores, how many more of the limiting ingredient would we need to buy? How many more S'mores could we make? *[We would need to buy two more chocolate bars. We could make six more S'mores.]*

   e. After the second purchase, what is the limiting ingredient, and what is left? *[The limiting ingredient is now the graham crackers. One half of a chocolate bar (6 squares), and 20 marshmallows would be left.]*

   If you feel brave enough and wealthy enough, you could bring the ingredients to class. Marshmallows can be toasted over a burner. (Marshmallow forks can be made from branches or wires. Low flames are recommended. If marshmallow drips inside the burner, a careful washing will remove it.)

   Safety Note: *Some teachers understandably feel that food should never be eaten in a laboratory setting. They point out that we can hardly emphasize this as a safety rule if broken occassionally.* This is an effective theoretical yield analogy without actually making and eating these treats in class.

3. **Making a BLT sandwich**: The number of sandwiches that can be made with each having three slices of tomato, four slices of bacon, one leaf of lettuce and two pieces of bread is limited by the number of slices of tomato, bacon and bread, and the number of lettuce leaves available. Students can determine the number of sandwiches (product) and the "leftovers" (remaining reactants) from a given quantity of each ingredient (reactants). (Other types of sandwiches could be used.)
4. **Clothes**: How many sets of clothes can be made from a given number of socks, shoes, slacks, shirts, *etc.*?
5. **Setting a table**: The number of place settings (of two spoons, one knife and two forks each) is limited by the number of knives, forks and spoons available. Pose problems such as : How many places can be set with 8 knives, 5 forks and 12 spoons? *[2]* What is the "limiting reactant"? *[Forks]* What flatware will be "leftover"? *[6 knives, 1 fork, 8 spoons]* This analogy was suggested by Robert C. Lamb.
6. **Vending machine**: If a vending machine will take several different denominations of coins, how many colas (or candy bars, chips, *etc.*) can be obtained from a given number of nickels, dimes and quarters? What coins remain? The cost of the item can be varied as well as the number of each type coin.
7. **Change machine**: If a given number of quarters is inserted in a change machine, how many dimes and nickels can be obtained, assuming 2 dimes and 1 nickel are given for each quarter inserted?

Pictures in the Mind

1. A balanced chemical equation can be visualized as a balanced see-saw or a balanced pair of scales (balance). Stoichiometry problems can be viewed as finding the balance point for the see-saw or scales by determining the masses of the remaining substances needed for balance when the mass of one substance is known.
2. Students should be reminded that you can't have more than 100% of anything. The percent yield cannot be over 100%! Visualize the theoretical yield as the maximum and the actual yield as something less than the maximum. Sometimes the percent yield is more than 100%. This could be due to impurities present contaminating the product or incorrect masses. Another possibility is that the proposed reaction is incorrect.

Other Activities

1. Batman and Robin in Chemistry: Stoichiometry problems based on comic strip (see *Appendix*).
2. Cork shooting contest: Students compete making the "best" cork shooting bottle, given certain parameters for quantities of sodium bicarbonate, acetic acid and bottle size (see *Appendix*).

TIPS FOR THE TEACHER

Language of Chemistry

stoichiometry relationships between quantities of substances (reactants and products) involved in a chemical reaction.

reactant substance that enters into chemical reaction, usually with one or more additional substances (also reactants). Reactants appear on the left in a chemical equation.

product substance formed as the result of a chemical reaction. Products appear on the right in a chemical equation.

Law of Conservation of Mass Mass (matter) is neither created nor destroyed in any chemical process. (In this module, the specific application is that mass is neither created nor destroyed in chemical reactions. The total mass of reactants is the same as the total mass of products in a chemical reaction. This law implies that the number of atoms, hence the moles, of each element is the same on both sides (reactants and products) of a chemical equation and that when an equation is balanced the number of each type atom is the same on each side of the equation.)

balanced equation represents a chemical reaction that conforms to the Law of Conservation of Mass. The molar amount of atoms of each element is the same on both sides (reactants and products) of a balanced chemical equation. Equations are balanced by placing a coefficient in front of a substance's formula to indicate the number of units of that substance required to bring the equation into balance.

mole ratio (also called stoichiometric ratio) fraction expressing the relationship of moles of one substance in a chemical reaction to moles of some other substance in the reaction. The mole ratio is derived from the coefficients in the chemical equation.

theoretical yield quantity of that product predicted from the stoichiometry of the reaction. (This statement assumes that the reaction goes to completion.)

actual yield quantity of product that is obtained when the reaction is carried out. The actual yield is never more than the theoretical yield and is frequently significantly less. The actual yield may be limited by the nature of the reaction, by the reaction conditions, by the inability to separate the desired product from remaining reactants or other products, or by the reactants producing "side products" by undergoing simultaneous competing reactions that do not yield the desired product.

percent yield fraction of the theoretical yield that is actually obtained, expressed as a percent.

$$\% \text{ Yield} = (\text{Actual yield}/\text{Theoretical yield}) \times 100$$

molarity (or **molar concentration**) moles of solute per liter of solution. This unit is used by chemists to express the concentration of solutions because the stoichiometry of chemical reactions is based on moles. If the concentration is expressed in molarity, the number of moles desired can be directly related to the volume (in liters) of the solution measured, facilitating stoichiometric calculations (see *Solutions* module).

Pattern Recognition

Coefficients in an equation relate moles of substances in a reaction. These coefficients are the numbers used in constructing the mole ratio for a stoichiometry problem.

Common Student Misconceptions

1. **"Subscripts can be changed to balance an equation."**

Only coefficients can be changed when balancing equations. Changing subscripts changes the nature of the substances, if the "new formulas" represent real substances at all! The reaction is not the same even if mass balance is obtained.

2. **"The number of reactant moles must be the same as the product moles in a balanced chemical equation."**

   This is generally not true because the atoms are arranged differently in the products and reactants. It is important to practice balancing equations where the number of moles of reactants and products differ.

3. **"The coefficients in the balanced equation relate the masses of the substances."**

   It is important to emphasize that stoichiometric coefficients relate *moles* of substances in the balanced equation. The masses of the substances will vary because molar masses for different substances are not equal. (See section on *Problem Solving* for suggestions on how to deal with this misconception.)

4. **"Changing the coefficient for a substance changes the molar mass."**

   Students may use the product of the molar mass and the coefficient as the molar mass. When the coefficient is also used in the mole ratio, the stoichiometry is incorrect.

5. **"Volumes of reactant solutions are related by the stoichiometric coefficients."**

   This may be a problem if laboratory activites and demonstrations involve equal volumes of all reactants or volumes proportional to stoichiometric coefficients. Take care in selecting laboratory activities and demonstrations so that the solution volumes used are not always in the ratio of the moles of reactants.

6. **"The molar mass for hydrogen, oxygen, chlorine, *etc.* is the same as the atomic mass."**

   In this module, the common diatomic gases should be reviewed. Students frequently use one-half the correct molar mass for hydrogen, oxygen, nitrogen and the halogens in stoichiometry problems because they assume all elements are monatomic.

Problem Solving

Balancing Equations

The following rules will provide guidance for students in balancing equations.

Five Commandments for Balancing Equations

1. Make a "score sheet" to keep track of elements in reactants and products. List elements in the order they appear in the reactants. Make an identical list for products. Record the number of atoms of each element for reactants and products. (If a polyatomic ion, such as NH_4^+ or NO_3^-, remains *unchanged* during the reaction, it can be treated in the same manner as an element.)
2. Fractional coefficients may be used. Coefficients can be converted to integers when the balancing is complete.
3. Use the following order for balancing the equation:
   a. Metallic elements
   b. Polyatomic ions that are *exactly the same* on both sides of the equation
   c. Nonmetallic elements that are *uncombined* on one side of the equation
   d. Remaining elements other than hydrogen or oxygen
   e. Hydrogen
   f. Oxygen

4. To remove fractional coefficients, multiply by an integer that will convert the fractions to integers. The integer should appear in the denominator of the fractional coefficient.
5. Using your original equation (*not the "score sheet"*), check that the equation is balanced (the same number of each atom type on both sides of the equation).

Solving Stoichiometry Problems

1. Emphasize the mole relationship in solving stoichiometry problems. Stoichiometry problems should be introduced by working mole-mole problems until the student is thoroughly familiar with the principle involved. The Mole A-Mole B conversion, and the mole ratio are the heart of stoichiometry problems.

   Then the mole-mass and mass-mole problem is introduced, again emphasizing that moles (not mass) of substances are related. Outlining the problem solution by making a "road map" is useful. For example, in the problem:

   "What mass of Al_2O_3 can be prepared by the reaction of 2.50 mol of Al according to the equation: $4\ Al + 3\ O_2 \rightarrow 2\ Al_2O_3$?"

   The "road map" would be:

   $$\text{Moles Al} \xrightarrow{\text{Mole ratio}} \text{Moles } Al_2O_3 \xrightarrow{\text{Molar mass } Al_2O_3} \text{Mass } Al_2O_3$$

   $$\text{Mole ratio} = \frac{2 \text{ mol } Al_2O_3}{4 \text{ mol Al}}$$

   The "road map" states steps involved in solving the problem. Start with the known (given) quantity and devise a plan to obtain the desired quantity. For each step, state the information needed to make the "conversion." Emphasize the centrality of the Mole A to Mole B conversion when working mass-mole and mole-mass problems.

   If sufficient practice is provided with these two types of stoichiometry problems, the transition to mass-mass problems (the most common and practical type problem) should be smoother, and students should have fewer problems. This type of problem adds an additional step to the "road map," but the step is the same as has already been encountered. There is a mass-mole conversion on "both ends" of the mole-mole conversion. Be sure to tell the students when to stop (when the desired unit is obtained).

2. Most texts use the factor-label method (dimensional analysis) for solving stoichiometry problems. Emphasize the importance of:
   a. Keeping the formula for the substance and units with each number.
   b. Outlining the "road map" as a guide to setting up the problem.
   c. All stoichiometry problems involve a mole ratio (to convert moles of one substance into moles of another substance).
   d. The molar mass relates mass to moles.
   e. The molarity relates moles to volume, in liters, of a solution.
   f. The mole ratio involves "counting numbers," and thus, does not limit the significant figures in the answer.
   g. The molar mass should be determined to at least as many significant figures as the given (or known) quantity in the problem.

h. After setting up the problem, check that units cancel to give the desired unit before manipulating the numbers in the problem. If the answer has incorrect units, the numerical part will not be correct either.

3. Most beginning students find it easier to understand limiting reactant problems if they calculate the yield of product from the given quantity of *each* reactant. They then compare the yield from each reactant. The maximum yield for the reaction (theoretical yield) will be the *smallest* of the calculated yields (see *Metaphors and Analogies*).

   Since the mass of product is usually desired, one of your answers (the theoretical yield) will be the answer to the problem. If the quantity of "leftover" reactants is desired, calculate the quantity of each reactant that will combine with the limiting reactant. Subtract the amount that reacts from the given amount to determine the quantity of "leftover" reactant.

4. It is important to generalize the approach to the different types of stoichiometry problems so that students begin to see the pattern in the solution:
   a. Given data about A, find Mole B.
   b. Relate Mole A to Mole B.
   c. From Mole B, find desired information about B (if answer has units other than Mole B).

   The mole ratio is the heart of stoichiometry. The previous approach will be used when gas reactions are studied and in the discussion of ionic equilibrium, electrochemistry, *etc.* It is central to all topics involving chemical reactions.

   Although stoichiometry problems involving heat may be discussed at this point or in the *Thermochemistry* module, reviewing stoichiometry at a later point in the course emphasizes the relationship of this key chemical concept to other topics.

Decision Making

It is important to stress the concept that "you can't get something from nothing" and "matter never disappears." This applies to chemistry (Law of Conservation of Mass) as well as to life itself. Pollution and waste disposal problems facing communities and nations are linked to stoichiometry. Students should realize that chemical reactions can be used to change the form of matter (*e.g.*, change a harmful or useless substance into something less harmful or even something useful), but it is impossible to make pollution or waste "disappear." Factors to consider, in addition to the feasibility of the process, are cost (economics), environmental effects of the clean-up (sometimes the by-products of clean-up are harmful), public perception of the problem, and the proposed solution.

Ask students to assume that they are managers in a chemical manufacturing plant. Discuss the factors to be considered in their responsibilities to their employer and society. Factors include alternative manufacturing processes (different chemical reactions with different starting materials to produce the same product—economics and stoichiometry are important issues here), source and consumption of raw materials (calculation of amounts of materials using stoichiometry is important here as well as economic and environmental considerations), waste disposal, pollution and public health issues (again economics and stoichiometry enter the picture). Students should be guided to see that there is no simple "right answer" here. Like other decisions in our lives, there are tradeoffs. As responsible citizens, we must make informed decisions concerning complex issues. The study of chemistry prepares us to make those decisions.

HISTORY: ON THE HUMAN SIDE

Scientists who contributed to our understanding of the quantitative relationships in chemical reactions include:

1. **Joseph Louis Gay-Lussac** (1788-1850, France). His experiments led him to propose in 1808 the Law of Combining Volumes, which states that the volumes of gases involved in a chemical reaction are in a small whole number ratio.
2. **Amedeo Avogadro** (1776-1856, Italy). In 1811, the Avogadro Hypothesis (equal volumes of gases at the same temperature and pressure contain equal numbers of molecules) was deduced from his observations of gas reactions and those reported by Gay-Lussac.
3. In the period 1949-1953, **E. Chargaff** and co-workers concluded that the mole ratio in DNA for adenine: thymine and cytosine: guanine was always 1:1, thus paving the way for the discovery of the precise structure of DNA by J. Watson and F. Crick.

HUMOR: ON THE FUN SIDE

1. Word Search (see *Appendix* for master copy)

A N G D P I U B Z U A C H E K E R
~~S T P I R C S B U S~~ V Y P D L H Q
T P Y V M B O E I F D C L O I Y U
N G V Z L O Q E B T R D M L M T P
E W Q X N S L Z N I P I G A I V N
I V P P D V ~~B A L A N C E D~~ T V L
C E L A Y K T S R Q X U F R I R X
I S K J I C C C T I X S N I N J W
F W G K A E W P O R T W M H G N Z
F N X E Y W X D K Y H Y O O H V P
E ~~P R O D U C T~~ C F N U C I T U H
O I O J B E N H D D ~~L A U T C A~~ T
C B J W F D P P D Y B B K L I D V

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Substance that normally appears on the left side of a chemical equation.
2. Substance that appears on the right side of a chemical reaction.
3. These are always conserved in a chemical reaction.
4. Adjective describing the yield of product obtained when a reaction is carried out.
5. Number of moles of solute per liter of solution.
6. Chemical equations are balanced with these numbers.
7. Numbers never changed when balancing a chemical equation.
8. Type of reactant present in smallest stoichiometric amount is said to be this.
9. If a chemical expression conforms to the Law of Conservation of Mass, it is said to be ____.
10. The ________ ratio is the heart of stoichiometry.

Answers: 1. REACTANT 2. PRODUCT 3. ATOMS 4. ACTUAL 5. MOLARITY 6. COEFFICIENTS 7. SUBSCRIPTS 8. LIMITING 9. BALANCED 10. MOLE

2. See cartoons at end of module.

MEDIA

1. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. *Stoichiometry*, by C. A. Wilkie. Saunders Computer Chemistry Apple IIe Software Package.
3. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).

   a. For the Apple II computer running on ProDOS: AR 301, AR 302.

   b. For the Apple II computer: AP 301, AP 304, AP 305, AP 307.

   c. For IBM PCs and PC-compatibles: PC 2201, PC 2301, PC 2302.
4. "The Atom" from the Ring of Truth series by Philip Morrison includes about a 3-minute segment on the reaction of different volume ratios of gases. This series appeared on PBS television during 1987. Videotapes are available from PBS Home Video (Phone 800-424-7963).
5. "Changes" from the Ring of Truth series by Philip Morrison deals with chemical reactions as well as heat change. This series appeared on PBS television during 1987. Videotapes are available from PBS Home Video (800-424-7963).
6. TV Ontario has a videotape series that contains short videos that could be used in this module. Free previews are available. TV Ontario Video, 143 West Franklin Street, Suite 206, Chapel Hill, NC 27516.
7. Doing Chemistry Videodiscs are available from the Education Division, American Chemical Society, 1155-16th Street, NW, Washington, DC 20036.
8. Software published by *JCE:* Software, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   a. *ChemCalc: A Scientific Calculator*, by Robert Allendoerfer. Vol. III B, No. 2, for IBM PS/2 PC-compatible computers.

   b. *REDOX: A Tutorial on REDuction / OXidation Equations*, by Derek Davenport, Paul Groves and Dale Jensen. Vol. III A, No. 1, for the Apple II computer.

   c. *Bomb Calorimeter Simulation*, by David J. Olney. Vol. III A, No. 2, for the Apple II computer.
9. Videodisc published by *JCE:* Software, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   "The Mole–From Counting to Weighing" and "Blowing Balloons with Moles of Hydrogen," two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

INSTRUMENTATION

The *Electrochemistry* module has information on electrolysis and construction of electrolysis apparatus.

Links/Connections

WITHIN CHEMISTRY

1. **Acids and Bases**. The reaction of acids and bases to produce water and a salt is called a neutralization reaction. The number of hydrogen ions (from the acid) neutralized in the reaction is equal to the number of hydroxide ions (from the base) that react. Complete neutralization occurs when all H^+ ions from the acid and all OH^- ions from the base react. Many neutralization reactions are carried out in solution. If the acid and base are both strong electrolytes, the net ionic equation for the reaction is:

   $$H^+(aq) + OH^-(aq) \rightarrow HOH(l)$$

   Titration is the procedure used to determine the concentration of a solution by reaction with another solution of known concentration or with a known mass of a solid. Titrations are frequently employed not only in the chemistry laboratory, but in industrial analytical laboratories (see *Acids and Bases* module).
2. **Chemical Equilibrium**. Writing a chemical equation often implies that reactants are completely consumed in forming products, but this is not always the case. Often there are significant amounts of both reactants and products when the reaction has apparently ceased (no measurable change in amounts of reactants and products can be detected). In such cases, the reaction has reached chemical equilibrium. The equilibrium concentrations of reactants and products can be used to determine the equilibrium constant, a quantity relating concentration of reactants and products present at equilibrium. For a reaction with a known equilibrium constant, the concentrations of all species present at equilibrium can be determined by knowing the amounts of substances added to the reaction vessel. Stoichiometry plays an important role in determining the concentrations of species present at equilibrium, since the coefficients in the balanced chemical equation are used in the mathematical relationship used to evaluate equilibrium concentrations (see *Equilibrium* module).
3. **Electrochemistry**. The amount of chemical change is related to the number of electrons transferred in an electrochemical reaction. The amount of chemical change in an electrolysis reaction is also related to current and time. For electrochemical reactions, stoichiometry problems may involve moles of electrons as well as reactants and products (see *Electrochemistry* module).
4. **Gas Stoichiometry**. Since one mole of an ideal gas occupies the same volume as a mole of another ideal gas at the same temperature and pressure, the volumes of reacting and product gases in a chemical reaction are in the same proportion as the stoichiometric coefficients. This relationship is referred to as Gay-Lussac's Law of Combining Volumes. (It should be noted that this relationship only holds for *gases*, not liquids or solids.) This relationship facilitates the determination of reactant and product volumes for gases (see *Gases* module).
5. **Net Ionic Equations**. Net ionic equations are important in solution chemistry because they show the "essential chemistry" of a reaction. In a net ionic equation, all ions not directly involved in the reaction ("spectator ions") are omitted. Frequently net ionic equations are more easily balanced than molecular equations. This is particularly true of oxidation-reduction reactions.

6. **Molarity and Solution Stoichiometry**. Many chemical reactions take place in solution. In these cases, concentration, rather than solute mass, is generally known for reactants. Since stoichiometric coefficients relate moles of reactants, the most useful concentration unit is molarity (moles solute per liter solution). Because net ionic equations are frequently used to describe reactions taking place in solution, it is important to be able to associate reactants, for which compound formulas are generally given, with the constituent ions, which appear in the net ionic equation. The stoichiometric coefficients obtained from the net ionic equation will be used in conjunction with molarity and/or molar mass in predicting quantities of reactants needed or products formed (see *Solutions* module).
7. **Thermochemistry**. Chemical reactions involve absorption or evolution of heat. The quantity of heat involved in a chemical reaction is directly related to the amounts of reactants and products (see *Thermochemistry* module).

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Chemical Engineering**. Commercial chemical manufacturing processes (such as the Contact process for sulfuric acid, the Ostwald process for nitric acid, and the Haber process for ammonia) involve chemical reactions. In selecting a reaction for industrial manufacture, consideration is given to yield as well as cost and availability of raw materials and adaptability to mass production (see *Industrial Inorganic Chemistry* module).
2. **Environmental Science**. An understanding of the environment requires knowledge of stoichiometry. For example, aerobic bacteria convert organic pollutants in streams to carbon dioxide and water. As the process occurs, dissolved oxygen is consumed, depleting the water of the oxygen required for the survival of marine life. Determination of the effects of dumping of organic chemicals (including sewage) into streams involves stoichiometry of the reactions involved.
3. **Geology**. Formation of stalactites and stalagmites involves the dissolution and reprecipitation of $CaCO_3$. Dissolved CO_2 causes limestone ($CaCO_3$) to dissolve according to the reaction:

   $$CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca^{2+}(aq) + 2HCO_3^-(aq)$$

   The reverse reaction results in reprecipitation of $CaCO_3$. (The reaction of $CaCO_3$ and $MgCO_3$ with dissolved CO_2 is also responsible for producing hardness in natural water supplies. See *Rocks, Minerals and Gems* and *Acids and Bases* modules.)
4. **Materials Science**. Relative proportions of starting materials and reaction conditions affect properties of polymeric materials. Most polymeric materials do not have fixed composition, that is, they contain molecules of varying molar mass but approximately the same proportions of constituent monomers. Properties (such as strength, resistance to chemical attack) of polymeric materials are correlated to reaction conditions to develop improved polymers. Glasses, ceramics, and metal alloys also are prepared with varying proportions of additives to alter the properties (see *Polymers* and *Materials Science* modules).
5. **Metallurgy**. The potential yield of metals from their ores can be determined from the reaction stoichiometry used in metallurgical processes. The ore's composition, as well as production costs and market demand determine the feasibility of development of mining operations of newly discovered mineral deposits.

6. **Biochemistry and Medicine**. Clinical laboratory scientists use stoichiometry in the study and analysis of body fluids. Metabolism involves the conversion of one substance into other substances. For example, in a known metabolic reaction, a product's (called a metabolite) concentration of a reaction in blood or urine can be used to determine the concentration of the original substance (see *Organic Chemistry, Chemistry in Medicine, Forensic Chemistry,* and *Enzymes: Biochemical Catalysts* modules).

TO THE CONTEMPORARY WORLD

Personal

Careers as research chemist, laboratory technician, medical technologist (clinical laboratory scientist), chemical manufacturer (including quality control as well as plant operations) involve the use of stoichiometry.

Community

1. **Field trips**: water treatment/sewage disposal plants, chemical industry, environmental testing laboratory, medical laboratory (clinical laboratory).
2. **Knowledgeable individuals**: chemical engineers, clinical laboratory scientists (medical technologists), environmental scientists/hygienists, industrial chemists, pharmacists, recycling engineers (resource management), research chemists, sanitation department chemists/water chemists, textile chemists

Societal (Science/Technology/Society; Current Events)

Stoichiometry is related to many important issues in everyday life and topics of current interest. A few of these are listed below.

1. **Comparison of Heat from Combustion of Various Fuels**. Fuels, such as coal, natural gas and petroleum, produce different amounts of heat energy upon combustion. Fuels, such as alcohols, that can be produced from renewable resources like grain or trees have been investigated as future energy sources. Electrical energy, solar energy, geothermal energy, and nuclear energy are also alternative energy sources. Comparison of the practicality of various energy sources for specific purposes involves many factors. One of these is combustion energy. Stoichiometry and thermochemistry provide the basis for the calculation of energy available from the fuel. Other factors are also important, however. Natural gas (methane) is a practical fuel for fleet vehicles (such as buses) that can return to a common refueling site each day but is not practical for private vehicles because of the complexity of refueling (see *Thermochemistry* module).
2. **Air Pollution and Fossil Fuel Combustion**. Burning coal or other fossil fuels containing sulfur or sulfur compounds generates sulfur dioxide, which produces acid rain (H_2SO_4) as shown in the series of equations:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$
$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$
$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Because of high temperatures reached in the internal combustion engine, nitrogen and oxygen from the air can react to form nitrogen oxides, also associated with air pollution. In addition to other effects on the atmosphere, nitrogen oxides also contribute to acid rain (HNO_3) as shown in the series of equations:

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$
$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$
$$2NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + HNO_2(aq)$$

Incomplete combustion of organic fuels (C_8H_{18}) can result in the formation of carbon monoxide, another air pollutant.

$$2C_8H_{18}(l) + 17O_2(g) \rightarrow 16CO(g) + 18H_2O(l)$$

3. **Ozone Depletion**. Freons (chlorofluorocarbons) released into the atmosphere from refrigerators, freezers, and air conditioners, result in photochemical reactions in the stratosphere. One reaction involves ozone destruction. Since ozone is responsible for absorbing ultraviolet light before it reaches the earth's surface where it increases the incidence of skin cancer in humans, ozone depletion has serious health implications (see *Photochemistry* module).

4. **Water Purification and Waste Water Treatment**. The first step in the purification of city water supplies involves allowing water to settle in large reservoirs. To facilitate this process, lime, and aluminum sulfate are added to produce gelatinous aluminum hydroxide, which settles slowly, carrying suspended materials, including bacteria, with it. Remaining bacteria are killed by addition of chlorine to the water.

$$3Ca(OH)_2 + Al_2(SO_4)_3 \rightarrow 2Al(OH)_3 + 3CaSO_4$$

This same process is used in the primary treatment of wastewater. Secondary and tertiary treatment of wastewater involves bacterial oxidation to convert soluble organic materials to substances that will settle, absorption of chemicals by activated carbon, and steam distillation of volatile organic pollutants.

Extensions

Batman and Robin in Chemistry (see *Appendix*)
Cork Shooting Contest (see *Appendix*)
Activity 1: Limiting Reagent for a Chemical Reaction in Solution in *Solutions* module

References

Module developed by Caroline L. Ayers, Philip Dail, and Carolyn Morse, the North Carolina team.

Acid-Base Reaction Saves Batman's Life. (1974). *Chemistry, 47*(9), 5.

The activity involving the Batman cartoon is based on this article and an adaptation supplied by Larry Duberich.

Cotton, F. A., Darlington, C. L., and Lynch,C. L. (1980). *Chemistry: An investigative approach*. Boston, MA: Houghton Mifflin.

This text contains many good student activities on stoichiometry.

Dail, P. (1984). *High school chemistry: A laboratory approach.* Raleigh, NC: Wake County Board of Education.

The percentage yield of a chemical reaction activity is adapted from a similar experiment given in this laboratory manual.

Doing Chemistry [Videodisc Program]. Washington, DC: Education Division, American Chemical Society.

Segments illustrate demonstrations, experiments, and techniques. Consult the *Teacher's Manual* supplied with the videodisc to locate segments.

Fenster, A. E., Harpp, D. N., Schwarcz, J. A., and Glanville, J. O. (1988). A well-known chemical demonstration to illustrate an unusual medical mystery. *Journal of Chemical Education, 65*, 621.

This article briefly describes a medical condition caused by failure of the body to reduce Fe^{3+} to Fe^{2+} in the blood and the treatment for the condition. An analogy is provided by the "blue bottle" demonstration, which also uses methylene blue in an oxidation-reduction reaction.

Gardner, M., Smith, P., and Heikkinen, H. (1990). *Laboratory assessment builds success*. Berkeley, CA: Lawrence Hall of Science, Institute for Chemical Education.

This laboratory manual contains tested student activities written in the same format as those in SourceBook.

Morrison, P., and Morrison, P. (1987). *The ring of truth: An inquiry into how we know what we know*. New York, NY: Random House.

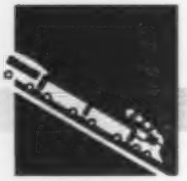
This book contains discussions about atoms, reactions, *etc.* that might be used in conjunction with this module.

Ramette, R. W. (1980). Exocharmic reactions. *Journal of Chemical Education, 57*, 68-69.

This article discusses experiments and demonstrations as motivational devices.

Szabadvary, F. (1972). Great moments in chemistry. Part VII. Chemistry learns to calculate (R. E. Oesper, Trans.). *Chemistry, 45*(4), 10-12.

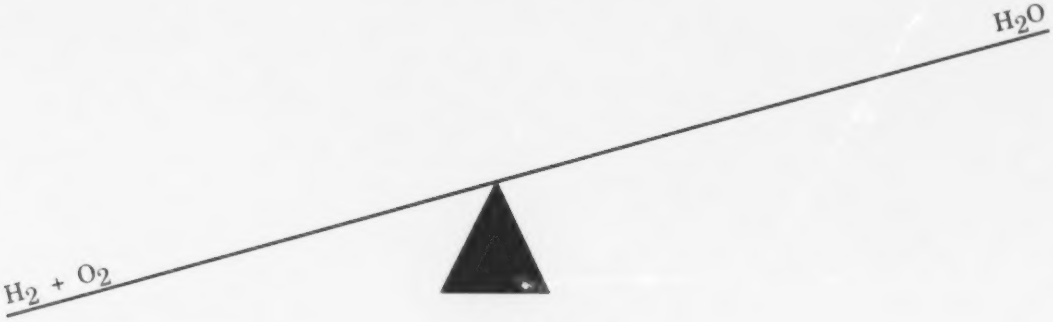
This is a play from a 15-part series originally presented over Hungarian radio and television. A reporter and a professor make a time reversal visit to Jeremias Benjamin Richter (1795) to discuss his early contributions to stoichiometry.

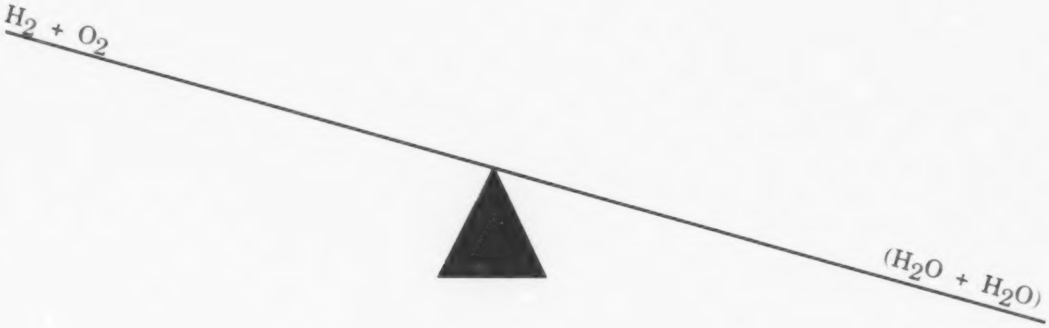


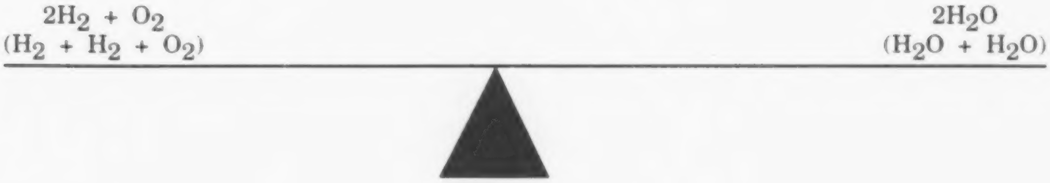
Appendix

- **Transparency Masters**
  1. Equation Balance
  2. Word Search
- **Humor**
  1. Batman and Robin in Chemistry
  2. Drawing of cylinder tank showing dimensions in Batman and Robin Chemistry
  3. Solutions to Batman and Robin in Chemistry
  4. Additional Activity: Cork Shooting Contest
  5. Cartoons

"Equation Balance"




$H_2 + O_2$
$(H_2O + H_2O)$


$2H_2 + O_2$
$(H_2 + H_2 + O_2)$
$2H_2O$
$(H_2O + H_2O)$

Word Search

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A N G D P I U B Z U A C H E K E R
S T P I R C S B U S V Y P D L H Q
T P Y V M B O E I F D C L O I Y U
N G V Z L O Q E B T R D M L M T P
E W Q X N S L Z N I P I G A I V N
I V P P D V B A L A N C E D T V L
C E L A Y K T S R Q X U F R I R X
I S K J I C C C T I X S N I N J W
F W G K A E W P O R T W M H G N Z
F N X E Y W X D K Y H Y O O H V P
E P R O D U C T C F N U C I T U H
O I O J B E N H D D L A U T C A T
C B J W F D P P D Y B B K L I D V
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Substance that normally appears on the left side of a chemical equation.
2. Substance that appears on the right side of a chemical reaction.
3. These are always conserved in a chemical reaction.
4. Adjective describing the yield of product obtained when a reaction is carried out.
5. Number of moles of solute per liter of solution.
6. Chemical equations are balanced with these numbers.
7. Numbers never changed when balancing a chemical equation.
8. Type of reactant present in smallest stoichiometric amount is said to be this.
9. If a chemical expression conforms to the Law of Conservation of Mass, it is said to be ____.
10. The ________ ratio is the heart of stoichiometry.

Batman and Robin in Chemistry

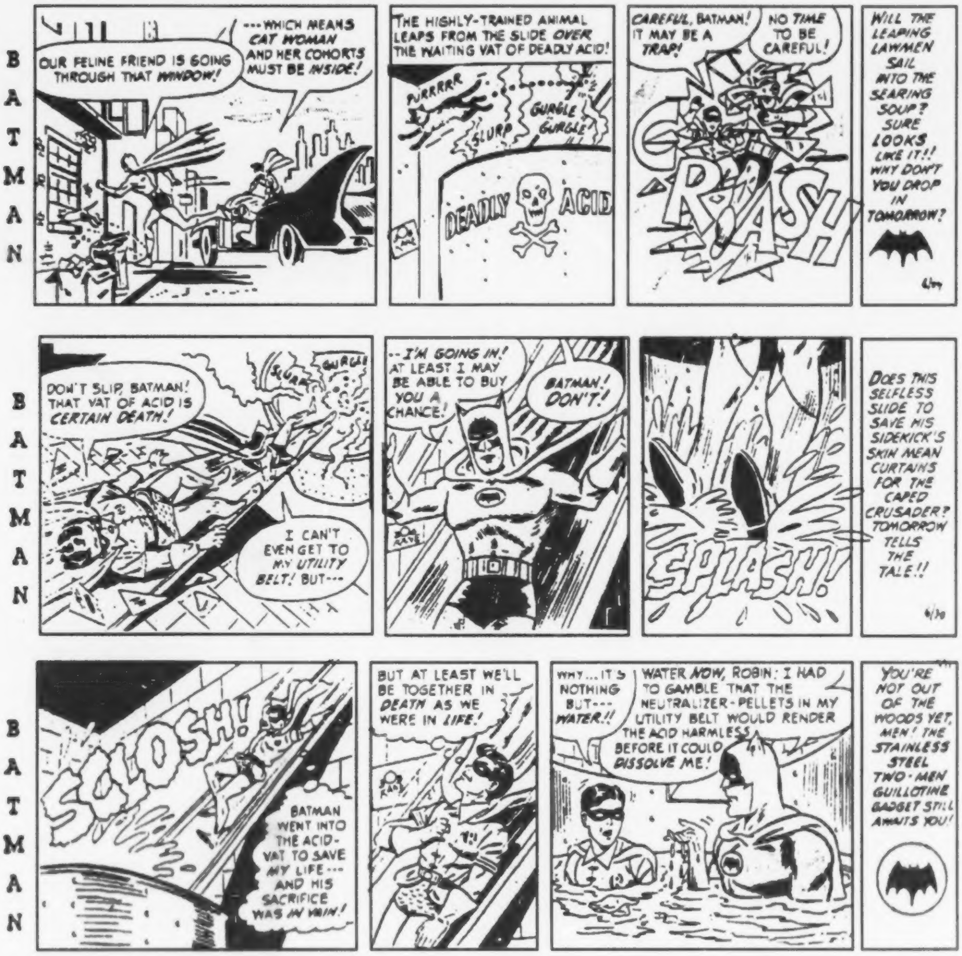
(Adapted from activity supplied by Larry Duberich. Used by permission of American Chemical Society. See first reference in *References* section.)

Batman uses pellets, which he does not identify, to neutralize the acid in the tank into which he falls. He does not give the size of the tank. Suppose the pellets were NaOH and the tank was 3.00 meters in diameter filled with 98% sulfuric acid by mass to a depth of 1.25 meters (see transparency).

Perform the following calculations and *show your work*.

1. Calculate the volume of acid (in cm^3) in the tank. (The volume of a cylinder is $\pi r^2 h$.)
2. Calculate the mass of acid in the tank, then find the number of moles of H_2SO_4 present. (Concentrated sulfuric acid is 98% (by mass) H_2SO_4 and has a density of 1.84 g/cm^3.)
3. One mole H_2SO_4 produces *two* moles of H^+ ions. With this in mind, write the balanced equation for the neutralization of NaOH and H_2SO_4. Determine the number of moles NaOH required to neutralize the acid in the tank. What *mass* of NaOH is required?
4. The heat of reaction for neutralization of a strong acid and a strong base is –57.3 kJ for each mole of H_2O formed. How much thermal energy (in kJ and J) was released in the neutralization reaction?
5. If the initial temperature of the acid bath was 20 °C, how much thermal energy would be needed to raise the temperature of the mixture in the tank (NaOH + H_2SO_4) to 100 °C? Assume the mass of the solution is the combined masses of the H_2SO_4 and NaOH. Also, assume the heat capacity of the liquid in the tank is 4.184 J/g.
6. Was all the generated thermal energy (Question 4) used to heat the liquid in the vat to 100°C (Question 5)? After water reaches the boiling point, an additional 2260 J/g is required to vaporize it. How much water was evaporated by the heat generated in the reaction of NaOH with H_2SO_4?

   (Ignore the boiling point elevation.)
7. In view of answers to Questions 3 and 6, how likely is it that the comic strip author passed high school chemistry?


B A T M A N
OUR FELINE FRIEND IS GOING THROUGH THAT WINDOW!
---WHICH MEANS CAT WOMAN AND HER COHORTS MUST BE INSIDE!
THE HIGHLY-TRAINED ANIMAL LEAPS FROM THE SLIDE OVER THE WAITING VAT OF DEADLY ACID!
PURRRRR
SLURP
GURGLE GURGLE
DEADLY ACID
CAREFUL, BATMAN! IT MAY BE A TRAP!
NO TIME TO BE CAREFUL!
CRASH
WILL THE LEAPING LAWMEN SAIL INTO THE SEARING SOUP? SURE LOOKS LIKE IT!! WHY DON'T YOU DROP IN TOMORROW?
B A T M A N
DON'T SLIP, BATMAN! THAT VAT OF ACID IS CERTAIN DEATH!
SLURP
GURGLE
I CAN'T EVEN GET TO MY UTILITY BELT! BUT---
--I'M GOING IN! AT LEAST I MAY BE ABLE TO BUY YOU A CHANCE!
BATMAN! DON'T!
SPLASH!
DOES THIS SELFLESS SLIDE TO SAVE HIS SIDEKICK'S SKIN MEAN CURTAINS FOR THE CAPED CRUSADER? TOMORROW TELLS THE TALE!!
6/30
B A T M A N
SCLOSH!
BATMAN WENT INTO THE ACID-VAT TO SAVE MY LIFE---AND HIS SACRIFICE WAS IN VAIN!
BUT AT LEAST WE'LL BE TOGETHER IN DEATH AS WE WERE IN LIFE!
WHY...IT'S NOTHING BUT---WATER.!!
WATER NOW, ROBIN: I HAD TO GAMBLE THAT THE NEUTRALIZER-PELLETS IN MY UTILITY BELT WOULD RENDER THE ACID HARMLESS BEFORE IT COULD DISSOLVE ME!
YOU'RE NOT OUT OF THE WOODS YET, MEN! THE STAINLESS STEEL TWO-MEN GUILLOTINE GADGET STILL AWAITS YOU!

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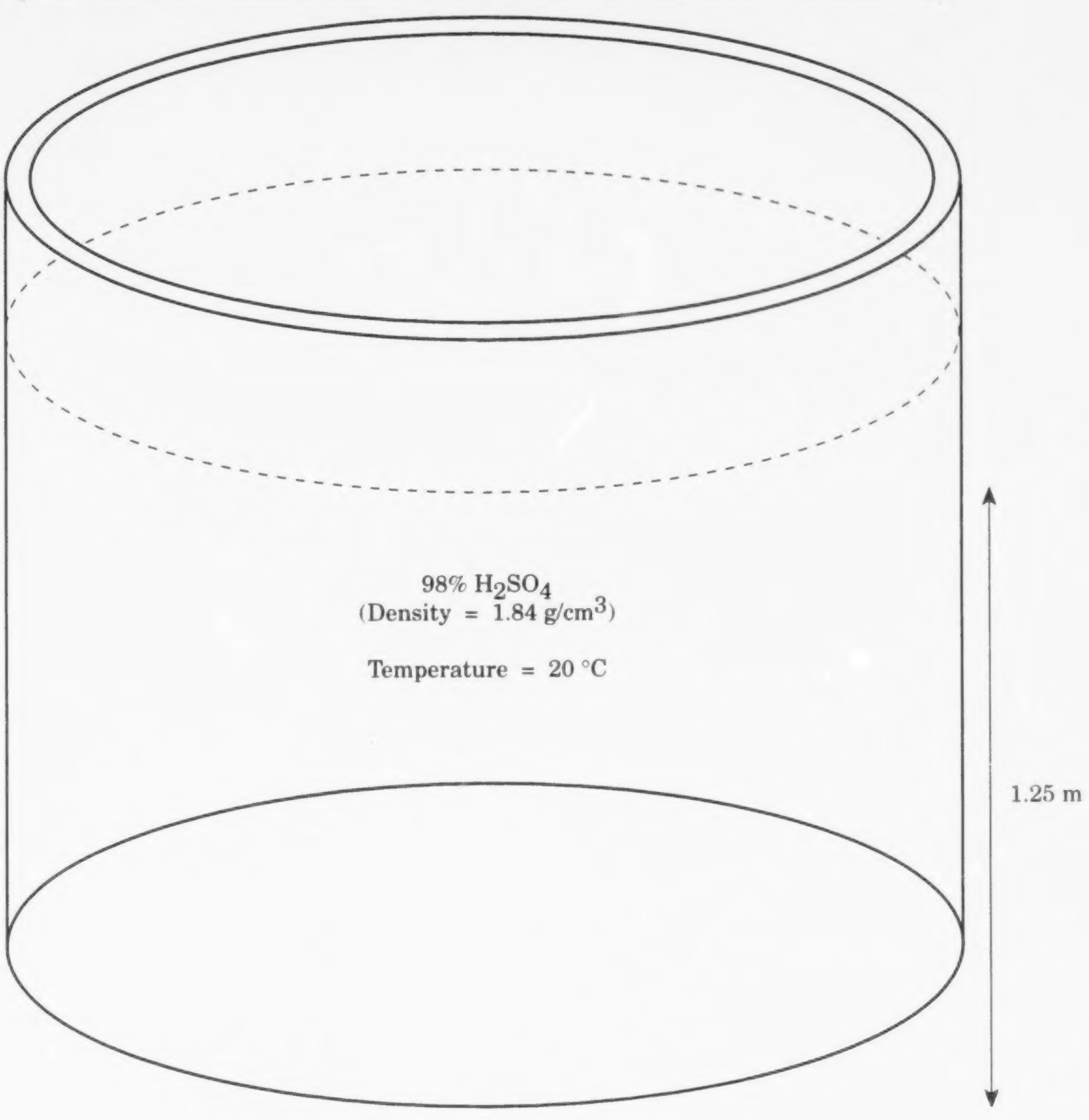
Cylindrical Tank


3.00 m
98% H_2SO_4
(Density = 1.84 g/cm^3)
Temperature = 20 °C
1.25 m

Solutions to Batman and Robin in Chemistry

1. Tank diameter = 3.00 m, $r = 1.50$ m, $r^2 = 2.25\ m^2$, $h = 1.25$ m

   $V = 8.84\ m^3 = 8.84 \times 10^6\ cm^3\ [1\ m^3 = (10^2\ cm^3)^3]$

2. $8.84 \times 10^6 cm^3 \times 1.84 \frac{g}{cm^3} = 1.63 \times 10^7 g$ acid

   $1.63 \times 10^7 g \text{ acid} \times \frac{98}{100} \times \frac{1\ mol}{98.1\ g} = 1.63 \times 10^5 mol\ H_2SO_4$

3. $H_2SO_4 + 2NaOH \rightarrow 2H_2O + 2Na^+ + SO_4^{2-}$

   $1.63 \times 10^5\ mol\ H_2SO_4 \times \frac{2\ mol\ NaOH}{1\ mol\ H_2SO_4} \times \frac{40.0\ g\ NaOH}{1\ mol\ NaOH} = 1.30 \times 10^7\ g\ NaOH$

   $1.30 \times 10^7\ g = 1.30 \times 10^4\ kg = 1.30 \times 10^1$ or 13.0 metric tonnes of NaOH

4. $1.63 \times 10^5\ mol\ H_2SO_4 \times \frac{2\ mol\ H_2O}{1\ mol\ H_2SO_4} \times \frac{57.3\ kJ}{1\ mol\ H_2O} = 1.87 \times 10^7\ kJ$

   $= 1.87 \times 10^{10}$ J released

5. Assume: solution mass = mass acid + mass NaOH

   1.63×10^7 g acid + 1.30×10^7 g NaOH = 2.93×10^7 g solution

   if initial temp = 20 °C, final temp = 100 °C

   heat absorbed = $2.93 \times 10^7\ g \times 80°C \times 4.184 \frac{J}{g\ °C} = 9.81 \times 10^9\ J$

6. Reaction produces: 18.7×10^9 J (1.87×10^{10} J)

   Heating to 100 °C uses 9.8×10^9 J. 18.7×10^9 J released – 9.8×10^9 J used = 8.9×10^9 J remain.

   $8.9 \times 10^9 J \times \frac{1\ g}{2260\ J} = 3.9 \times 10^6$ g solution are vaporized

   $\frac{3.9 \times 10^6 g}{2.93 \times 10^7 g}$ = 13% of the solution in the tank is vaporized

   (Note: The boiling point of a solution is greater than that of the solvent. This is ignored in these calculations.)

7. 13 metric tonnes of NaOH sounds like a serious Bat–Hernia. Furthermore, the Dynamic Duo were certainly cooked!

Cork Shooting Contest Rules and Regulations

(Adapted from activity submitted by John Garner)

Purpose

To determine what method of mixing sodium bicarbonate and acetic acid will shoot the cork the farthest.

Reaction

$NaHCO_3(s) + HC_2H_3O_2(aq) \rightarrow CO_2(g) + H_2O(l) + NaC_2H_3O_2(aq)$

Materials

Cork
Plastic Pop Bottles
Protractor
Tape measure
Baking Soda (Sodium bicarbonate, $NaHCO_3$)
Vinegar (5% acetic acid, $HC_2H_3O_2$)

Procedure

This is a student project involving concepts of chemistry and physics. *Students should observe standard safety procedures.* Design the procedure such that the optimum amounts of baking soda and vinegar are used to shoot the cork the maximum distance. Additionally, consider the bottle size and angle of tilt to maximize the cork's flight distance.

Safety

Use a plastic bottle. A glass bottle might explode during the reaction creating a safety hazard. (If a glass bottle must be used, it should be wrapped in heavy tape to prevent shattering in case of explosion.)

Rules

1. Must be done outside.
2. No more than two tablespoons baking soda ($NaHCO_3$).
3. No more than two tablespoons of vinegar ($HC_2H_3O_2$).
4. Cork must be decorated and named.
5. Data from 10 trials must be turned in before you can enter the contest.
6. Use 6, 10, 12, 16 or 32 ounce bottles only.
7. Bottle must be held on the ground.

Variables

1. Size of bottle
2. Quantity of vinegar
3. Angle of bottle
4. Quantity of baking soda
5. Design of cork
6. Mass of cork
7. Extent of shaking of bottle
8. Position of cork in the bottle

Information to Include in the Write-up

1. What makes the cork fly out of the bottle?
2. How many grams are in a tablespoon of vinegar? How many grams are in a tablespoon of baking soda?
3. Complete the tables of your trials at home and the trials of all your classmates from school during the competition at school.
4. What problems did you encounter and how did you solve them?
5. From all the data predict what the best variable combination is to get the longest distance.
6. Make a graph of distance *vs.* bottle size.
7. Make a graph of distance *vs.* amount of vinegar.
8. Make a graph of distance *vs.* amount of baking soda.

Report

1. Measurements: Distance cork travels
2. Data and Calculations: Questions 3, 6, 7, 8
3. Answer: Questions 1, 2, 4, 5

Date:

Name:

Home Laboratory - Cork Shooting Contest

Personal Data Sheet

To earn points, this worksheet may be turned in by:

__________________+ 5 __________________ + 3 __________________ + 1

These trials must be completed at home and turned in before: ______________

Each student should run at least 10 trials at home and record all the data below.

Name:	Name:
Design of Cork A	Design of Cork B

Trial	Cork Name	Volume Acid	Volume Soda	Type Bottle	Size Bottle	Angle of Bottle	Distance Traveled
1							
2							
3							
4							
5							
6							
7							
8							
9							
10							

This information should be included in the data and calculation section of your formal laboratory write-up.

Date:

Class Data Sheet

Weather:

Trial	Cork Name	Volume Acid	Volume Soda	Type Bottle	Size Bottle	Angle of Bottle	Distance Traveled
1							
2							
3							
4							
5							
6							
7							
8							
9							
10							
11							
12							
13							
14							
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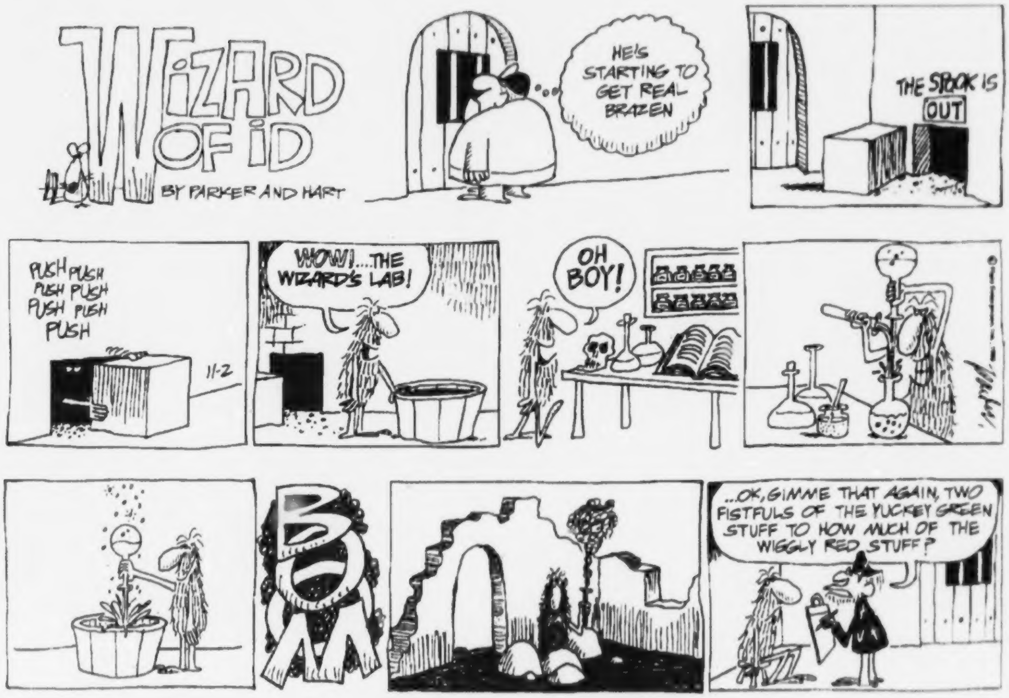
16							
17							
18							
19							
20							


WIZARD OF ID
BY PARKER AND HART
HE'S STARTING TO GET REAL BRAZEN
THE SPOOK IS OUT
PUSH PUSH PUSH PUSH PUSH PUSH PUSH
11-2
WOW!...THE WIZARD'S LAB!
OH BOY!
BOOM
...OK, GIMME THAT AGAIN, TWO FISTFULS OF THE YUCKEY GREEN STUFF TO HOW MUCH OF THE WIGGLY RED STUFF?

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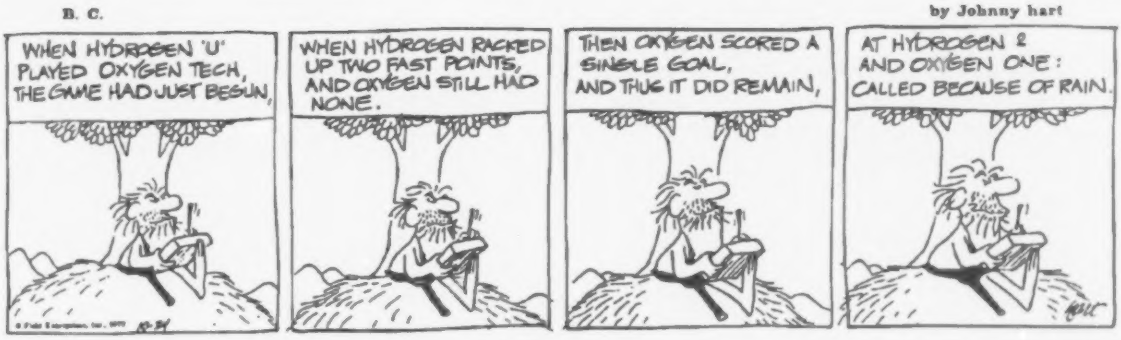
B. C.
by Johnny hart
WHEN HYDROGEN 'U' PLAYED OXYGEN TECH, THE GAME HAD JUST BEGUN,
WHEN HYDROGEN RACKED UP TWO FAST POINTS, AND OXYGEN STILL HAD NONE.
THEN OXYGEN SCORED A SINGLE GOAL, AND THUS IT DID REMAIN,
AT HYDROGEN 2 AND OXYGEN ONE: CALLED BECAUSE OF RAIN.
hart

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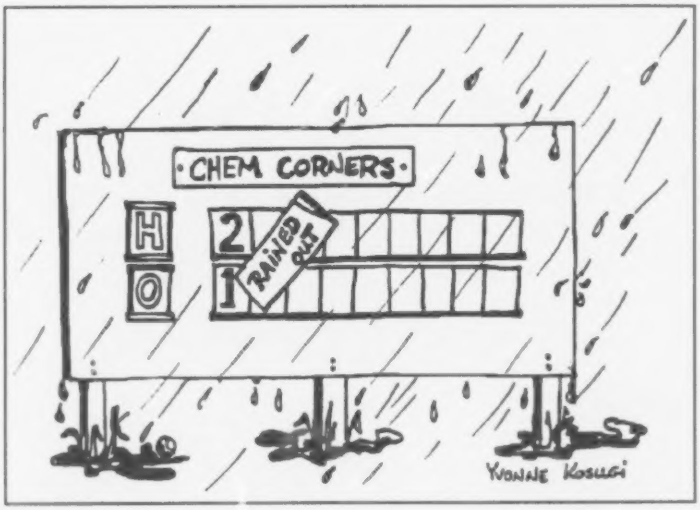
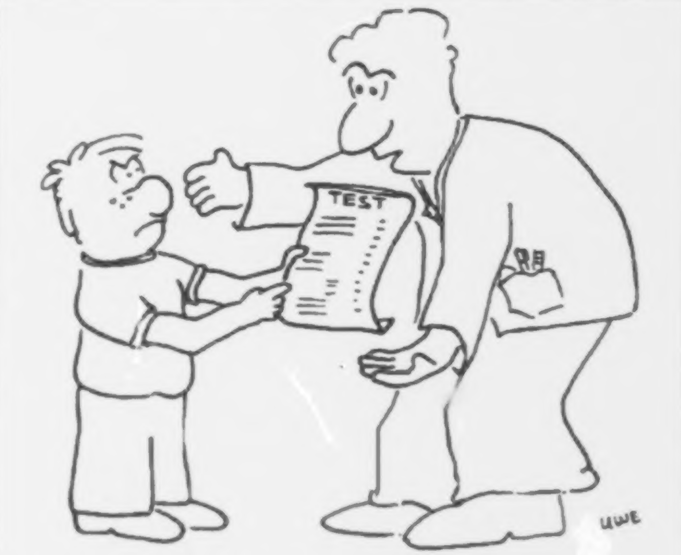

CHEM CORNERS
H
O
2
1
RAINED OUT
YVONNE KOSUGI

CHEM 13 NEWS, October 1984, p. 13. Reprinted with permission.



THE LAW OF MULTIPLE PROPORTIONS DOES NOT MEAN 3 ANSWERS WITH DIFFERENT PROPORTIONS !!!

CHEM 13 NEWS, January 1980, p. 13. Reprinted with permission.

UWE
THE REACTION ISN'T STOICHIOMETRIC
BUT IT LOOKS HUNGRY!!!

CHEM 13 NEWS. Reprinted with permission.

A SourceBook Module

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THERMOCHEMISTRY

ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

You feel cold before you dry off after a swim or shower because each mole of water on your skin requires approximately 40 kJ (10 kcal) of heat to evaporate, and this heat is taken from your body. The chemical cold packs used to treat sports injuries cool their surroundings because the chemical reaction absorbs heat from the water when the chemicals in the packs dissolve (go into solution). In contrast, concentrated sulfuric acid releases large quantities of heat when added to water. Chemical hot packs provide heat in much the same way—except that they do not contain sulfuric acid!

You are kept warm in winter by the heat released when fuels such as coal, oil, or natural gas are burned. Heat released from nuclear fission reactions in reactor cores is used to generate electricity by nuclear power plants, while the sun's energy comes from the nuclear fusion reactions deep in its interior. Our Earth receives only 1 part in 2.4 billion parts of the sun's energy (1.7×10^{17} watts out of 4×10^{26} watts). It would require 115 million nuclear (fission) power plants to generate this same quantity of energy. It takes only three days for the Earth to receive from the sun a quantity of energy equivalent to all fossil fuel reserves on Earth. The United States, which has over 40% of the world's coal reserves, burns about 700 million tons of coal each year to generate over half of our needed electrical energy. It is estimated that, by the year 2005, it will take as much energy to produce a barrel of oil as we would obtain by burning that quantity of oil as a fuel. Future energy sources include solar energy, biomass, wind power, hydroelectric, geothermal, tidal, fuel cells, and nuclear fission and fusion reactions.

Different substances have different quantities of energy stored in them as potential energy. When substances undergo physical or chemical change, energy is released or absorbed as heat and other forms of energy. Energy is conserved during such changes, which allows chemists to predict the quantity of energy released or absorbed. The quantity of energy released or absorbed can be measured indirectly using the technique known as calorimetry, which stems from the old definition of the unit of heat energy known as the calorie. Since one calorie (cal) is the quantity of heat needed to raise the temperature of one gram of water just one degree Celsius, calorimetry involves measuring the temperature change of a known mass of water accompanying the studied change. The IUPAC unit of heat energy, the joule (J), is the preferred unit. 1 cal equals 4.184 J, or 1 J equals 0.2390 cal.

If a water calorimeter is used and the water temperature drops during the change being measured, heat energy must have been absorbed from the water, and the change is said to be **endothermic**. If water temperature rises, heat was released to the water, and the change is **exothermic**.

Whenever chemical equations representing two or more changes can be algebraically combined to yield a new equation, the algebraic sum of the heats accompanying these changes is found to equal the quantity of heat associated with the net change. This regularity, known as **Hess's Law**, permits calculation of some heats of reaction without measuring them directly.

A table of standard heats of formation of compounds from their elements provides another tool students can use to make the same calculated prediction of heats of reaction.

PLACE IN THE CURRICULUM

Thermochemistry fits into the curriculum in several places. It is related to stoichiometry (heats of reaction), chemical bonding (bond energies), phase changes (heats of vaporization and fusion), solution chemistry (heats of solution and hydration), and nuclear chemistry (fission and fusion). It should be taught prior to the discussion of rates of reaction, since activation energies for endothermic and exothermic reactions are related to both chemical kinetics and equilibrium.

CENTRAL CONCEPTS

1. A **system** is that portion of the physical universe under study. Other portions are termed the **surroundings** or the environment.
2. **Heat** is a form of energy that flows across a conducting boundary from the system at a higher temperature to the system at a lower temperature (from a hot body to a cold body).
3. When chemical and/or physical changes occur, they involve definite amounts of matter accompanied by specific quantities of heat energy.
4. **Endothermic** processes are those in which heat is absorbed by the system from the surroundings; **exothermic** processes release heat from the system to the surroundings. Processes in which no release or absorption of heat occurs are termed **adiabatic** changes.
5. A **joule** (J) is the energy that a 1-kg mass gains when it is acted upon by a force of 1 N (newton) through a distance of 1 m. The older unit, the calorie (cal), is equivalent to 4.184 J.
6. **Heat capacity** of a material is defined as the quantity of heat required to raise the temperature of a substance by one Kelvin. (1 °C = 1 K) The specific heat capacity is the heat capacity per gram of substance. Common examples of specific heat capacity ($J\ K^{-1}g^{-1}$) are: water = 4.18; copper = 0.38; nitrogen = 1.04; and ammonia = 2.06. The molar heat capacity is the heat capacity per mole of substance.
7. **Calorimetry** is an experimental procedure in which the quantity of heat released or absorbed during a change is indirectly measured, at constant volume or at constant pressure, for a known quantity of matter. From the definition of a calorie, it follows that this quantity of heat for a water calorimeter can be calculated using the relation

Quantity heat	=	Mass water	x	Temp. change water	x	Specific heat water
J (or cal)		(g)		(K or °C)		($JK^{-1}g^{-1}$ or cal/g °C)
Q		m		ΔT		C

   $Q = m \times \Delta T \times C$
8. The heat involved in a physical or chemical change occurring at constant pressure is known as the **enthalpy change**, given the symbol ΔH.
9. Tables listing the quantities of heat accompanying many chemical changes have been compiled, in units of both joules and calories. When such changes involve the formation of one mole of a chemical compound from its elements in their normal (standard) states, the values are known as standard molar heats of formation, ΔH_f^0.

10. The heat energy term associated with a chemical or phase change can be included in the corresponding equation. Heat can be treated as a product or a reactant in an equation. If heat is absorbed during the change, the heat term is treated as a reactant, and placed to the left of the arrow in the equation. If heat is released, this term is treated as a product, and is placed to the right of the arrow, with the substances produced in the reaction. Alternatively, the heat energy term can be listed apart from the equation, at the end, as a ΔH notation, where a + sign indicates an endothermic reaction and a − sign indicates an exothermic reaction.

11. The first law of thermodynamics sets down the empirical principle that the internal energy change (ΔE) in a system equals the sum of the heat (q) and work (w) in that system ($\Delta E = q + w$). The internal energy is the sum of the translational, rotational, vibrational, electronic, nuclear, and zero point energies ($E = E_{trans} + E_{rot} + E_{vib} + E_{elec} + E_{nuclear} + E_{zero\ point}$).

12. The total energy of a system is composed of kinetic energy and potential energy. **Kinetic energy**, energy of motion, is directly related to temperature. The higher the temperature, the greater the kinetic energy of the system. **Potential energy**, energy of position, is associated with attractions and repulsions among (and within) the molecules, atoms or ions present in the system. Different substances are assumed to contain different quantities of stored energy (potential energy). When one substance changes to another, with a new potential energy content, energy is either released or absorbed as heat energy. The total kinetic energy of a polyatomic system consists of three types: translational, rotational, and vibrational. Solids exhibit vibrational energy only, while liquids and gases can exhibit all three modes of motion.

13. Phase changes, such as melting and evaporation, involve energies of 2-10 kJ/mol. Chemical changes involve ten to a hundred times more energy than phase changes. Nuclear changes involve energies millions of times greater than chemical changes.

14. The assumption that total energy is conserved during a chemical or physical change permits the calculation of the change in stored energy, also known as the change in heat content (ΔH), when the change is measured in a calorimeter. If one mole of A absorbs 10,000 J when changing to one mole of B, for example, $\Delta H = +10$ kJ per mole of B formed. One mole of B would have 10 kJ more potential energy than did the one mole of A. If 10,000 J had been released in the process, then $\Delta H = -10$ kJ/mol of B. Mole for mole, in this latter case, B would have 10 kJ less potential energy than A.

15. Whenever equations representing two or more changes can be combined algebraically to give the equation of another, different change, the algebraic sum of the heats accompanying these changes is found to equal the quantity of heat associated with the net change. This is known as **Hess's Law**.

16. From the concept that heat (enthalpy) of reaction is ΔH (where "delta" means a change in some fundamental property) one can calculate the heat of reaction for a given equation using standard molar heats of formation and the following algebraic relationship:

$$\Delta H = \Sigma\, \Delta H_f^o(\text{products}) - \Sigma\, \Delta H_f^o(\text{reactants})$$

17. Enthalpy diagrams based on reaction mechanisms can be used to gain understanding about the reasons why certain physical and chemical changes occur.

18. Bond dissociation energies are useful in allowing one to predict the heat of reaction for specific chemical reactions.
19. ***Optional:*** When the atoms, molecules, or ions in a system become more disordered (more randomly distributed) during any change, the property known as entropy (S) is said to increase. The change in entropy is symbolized as ΔS. Changes will occur spontaneously when the difference between the enthalpy change (ΔH) and the absolute temperature (T), multiplied by the entropy change (ΔS) gives a negative value. This difference is known as the free energy change and is symbolized by ΔG. Thus, $\Delta G = \Delta H - T\Delta S$. Reactions occur spontaneously (proceed more in the forward direction) when ΔG is negative; reactions are not spontaneous (proceed more in the reverse direction) when ΔG is positive. When $\Delta G = 0$, the system is at equilibrium.

RELATED CONCEPTS

1. Chemical equations (*Stoichiometry* module)
2. Mole concept and stoichiometry (*The Mole* and *Stoichiometry* modules)
3. Solution concentration and molarity (*Solutions* module)
4. Kinetic molecular theory of gases (*Gases* module)
5. Molecular shapes and structure (*Molecular Geometry* module)
6. Difference between temperature and heat (*Gases* module)

RELATED SKILLS

1. Manipulative skills: Use of graduated cylinders, balances, thermometers.
2. Mathematical and conceptual skills:
   a. Problem solving techniques (*Introductory* and *Stoichiometry* modules)
   b. Dimensional analysis and unit conversions (*Introductory* module)
   c. Estimating answers
   d. Simple algebra—combining several equations

PERFORMANCE OBJECTIVES

After completing their study of thermochemistry, students should be able to:

1. differentiate between a system and its surroundings.
2. determine the direction of heat flow in a system that undergoes a temperature change.
3. differentiate among endothermic, exothermic, and adiabatic changes.
4. define the joule, the calorie.
5. differentiate among heat capacity, specific heat capacity, and molar heat capacity.
6. given an initial temperature, final temperature, and mass of water involved in a water calorimetry experiment,
   a. determine if heat was absorbed or evolved.
   b. calculate the heat change in that system.
7. write the heat term involved in a reaction including it as part of the chemical equation and using ΔH notation.

8. discuss an exothermic (or endothermic) reaction in terms of heat changes and potential energy differences.
9. differentiate among gases, liquids, and solids in terms of the three types of kinetic energy—vibrational, rotational, and translational.
10. discuss the total energy of any system.
11. compare on a relative scale the energies involved in physical, chemical, and nuclear changes.
12. using a table of standard heats of formation, or known heats of reaction, use Hess's Law to calculate the heat of reaction for a reaction with unknown ΔH.
13. given a table of standard heats of formation, calculate ΔH for a given reaction.
14. determine the nature of a reaction, *i.e.*, exo- or endothermicity, from its enthalpy diagram.
15. construct an enthalpy diagram given a description of the energetics of the reaction.
16. predict the heat of reaction when bond dissociation energies are given.
17. ***Optional***: Discuss entropy as it relates to a specific chemical or physical change, *e.g.*, a phase change from liquid to gas.
18. ***Optional***: Given a table of standard heats of formation and standard entropy values, calculate ΔG for a given reaction.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: The Heat of Combustion of a Candle

Introduction

A candle gives off light and heat energy as it burns. You might wonder, how much heat. To find out, you will use the definition of the unit of heat energy, the joule, and measure the temperature change produced in a known mass of water by a burning candle. You will also determine the mass (in g) of candle that has burned, so your results can be compared to those of your classmates in a meaningful way. Your final calculated result, in J/g, is the heat of combustion.

Purpose

To determine the quantity of heat given off when a candle burns (combusts).

Safety

1. Wear protective goggles throughout the laboratory activity.
2. A candle flame is hot, and the can holding the water (as well as the water) may also get hot, so exercise caution to avoid burning your hands.

Procedure

To make optimal use of your time, you should collect the data in the order that follows:

1. Determine the mass of the candle to the nearest 0.01 g.
2. Measure the temperature (°C) of cold water provided by your teacher in your "calorimeter" (tin can).
3. Record these two measurements, then immediately light the candle and place it under the can of cold water (see Figure 1).
4. Record room temperature (°C).
5. Stir the water gently with a stirring rod, not the thermometer, and allow the candle flame to heat the water as far above room temperature as it was below room temperature at the start. (For example, if the water was at 5 °C when heating began, and room temperature is 25 °C, heat the water to a temperature of 45 °C before blowing out the candle.) Keep stirring gently and record the highest temperature reached by the water as its final temperature.

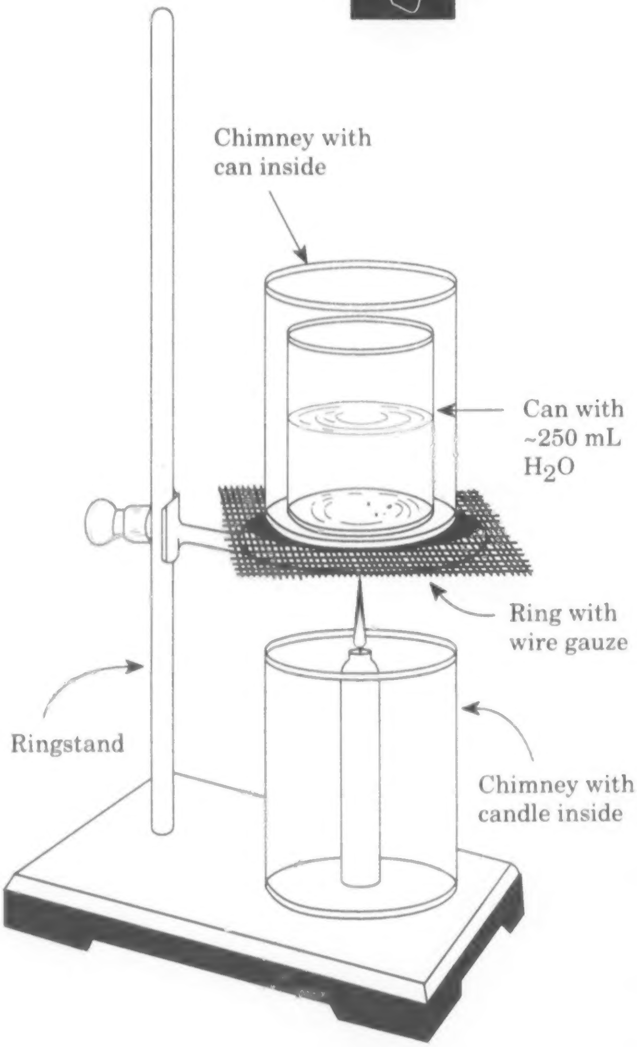


Figure 1. Apparatus for **Activity 1.**

6. Determine mass of the can with its heated water, but only to the nearest gram.
7. Discard the water into the sink, dry the can, and determine the mass of the empty (and dry) can, again only to the nearest gram.
8. Finally, find the mass of the candle, to the nearest 0.01 g. Be sure to include any hardened drippings.
9. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

1. Organize your six measurements into a neat, concise data table, assigning each a capital letter (A,B,C,D,E,F).
2. Make a table of calculated results, and indicate how each calculation was made, using letters. For example, if the mass of the candle before burning is represented by "A," and its mass after burning by "B," then the mass of candle burned, say "L" equals A – B. Place the numerical result of this calculation under its symbol. Be sure to include the uncertainties, both in the "Raw Data" (your measurements) and your calculated results.
3. Think about the procedural directions you were given—the "strategy" of the activity. Why were you given cold water to heat instead of room temperature water? Why heat it up as far above room temperature as it initially was below that temperature?
4. Now think about uncertainties, both in your original six measurements and the calculated results generated from them. How many places are justified in your values for the heat absorbed by the water, and for the heat of combustion of the candle?
5. Finally, think about sources of error. How did your teacher attempt to minimize them? Will your experimental heat of combustion value due to these errors be too high or too low (relative to a standard heat of combustion)? Try to support your answer.

Data Tables

Student Results

Raw Data	**Candle (± 0.01 g)**		**Tin can (± 1 g)**		**Temp. (± 0.5 °C)**	
	Before	*After*	*Empty*	*+ Water*	*Before*	*After*
	A	B	C	D	E	F

Calculated Results

	Mass of candle burned (g)	**Mass of water warmed (g)**	**Temp. change of water (°C)**	**Heat absorbed by water (J)**	**Heat of combustion of candle wax (J/g)**
	L = A – B	M = D – C	N = F – E	O = M x N x 4.184 J/g°C	$P = \frac{O}{L}$
Value					
Uncertainty					
% Uncertainty					

Implications and Applications

1. Complete the following table of information for other fuels:

Fuel	Heat of combustion (kJ/g)	Molar heat of combustion (kJ/mol)	Temperature rise for 1 L water per 1 g fuel (°C/L/g)
C	32.8	394	7.84
CH_4	55.7		13.3
C_2H_6		1560	12.4
C_3H_8	52.8		
C_2H_5OH		1390	
$C_6H_{12}O_6$			3.70
H_2		287	
U-235	8.00×10^7		

2. Calculate the relative value of uranium as a fuel compared to coal (coal is mostly elemental carbon, C). (Use a ratio.)
3. Discuss the advantages and disadvantages of using hydrogen gas as a fuel in cars.
4. Discuss the advantages and disadvantages of using uranium as a fuel in nuclear reactors.

References

Similar activities are presented in:

Merrill *et al.* (1987). *Experimental foundations lab manual* (pp. 39-40). Prentice-Hall.

American Chemical Society. (1993). *ChemCom: Chemistry in the community* (pp. 186-190). Kendall-Hunt.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: The Heat of Combustion of a Candle

Major Chemical Concept

This laboratory exercise will allow students to conduct a calorimetry experiment, where they must decide what measurements they need to make, given the definition of a joule, and that:

Heat (J) = Mass water (g) x C (4.18 J/g °C) x Temperature change of water (°C)

You should emphasize quantitative results and the accompanying uncertainties, given the very crude calorimeter used.

Level

All college preparatory classes of chemistry (and physics).

Expected Student Background

Students are expected to have been exposed to the concept of calorimetric measurement and uncertainty in measured data as well as in calculated results derived from these data.

Time

One or two class periods will be needed for this activity. The length of time available will determine how many different trials each student group will be assigned.

Safety

Read the *Safety Considerations* in the *Student Version*. Students should be cautioned about the hazards of an open candle flame and, eventually, hot water and hot tin can, at their laboratory bench. No other specific cautions are needed for this activity, except to follow the general safety guidelines already established for other activities.

Materials (For 24 students working in pairs)

Nonconsumables

12 Thermometers, –10-110°C
12 Tin or aluminum cans, approximately 300 mL capacity
24 Larger cans, at least 1 L capacity (for "chimney")
12 Ring stands and rings
1 Laboratory balance, accurate to 0.01 g
1 Balance (0.1 or 1 g)

Consumables

12 Paraffin candles, 2-3 cm diameter x 18-20 cm length
Matches
Ice cubes for cooling water bath (*OPTIONAL:* Only used if teacher does not supply 1-2 L cold water.)

Advance Preparation

If possible, have 2 L cold water ready as a source to be poured into the smaller tin can for each group of two students. Each larger can (chimney) should have its top and bottom removed with a can opener, and several smaller cuts should be made at one end to allow air to enter when the larger can is placed over the candle apparatus in order to minimize wind disturbances, *etc.*

It is suggested that students also enter their measurements on a master sheet for later analysis and comparisons. Alternatively, a computer program or spreadsheet can be used to gather student data as well as the calculated heat of combustion. If a computer program or spreadsheet is to be used, have the computer ready for data entry by students.

Pre-Laboratory Discussion

Encourage students to list essential measurements needed, and then talk about the order in which they should be made (see *Student Version*). Show the suggested set-up, including the use of the larger can to "protect" the candle and smaller can from room air currents. Point out the need to be able to calculate the mass of candle burned, giving the definition of heat of combustion (J/g). Stress the importance of collecting all drippings from the candle, and determining the mass of these along with the rest of the still unburned candle at the end of the activity. You may want to instruct students to place the candle in a small baking cup or on a tin can lid to contain the drippings, which, if used, must be part of the candle's initial and final masses. Explain that it is essential that students quickly light the candle and place it so it begins to heat the can and cold water as soon as the temperature of the cold water is measured.

If a computer program or spreadsheet is to be used to collect data, explain how the data should be entered by students.

Encourage students to organize their data and calculated results so they know exactly how to use the data to obtain desired results. One way is to assign each measurement a letter and use these letters to indicate the calculations, as shown in the data tables.

Uncertainties need to be stressed so that students do not keep more significant figures in the calculated results than they should (see *Introductory* module).

Alternatively, you may choose to use the rule of significant figures (see *Introductory* module). If you choose this method, you will obtain a product (quotient) that has no expressed uncertainty; the implied uncertainty lies in the last significant figure that is allowed to be written. In the case of heat absorbed by water (O), 34,000 J, the calculated value (33,941 J) comes from the product of the specific heat of water, 4.184 J $K^{-1}g^{-1}$, mass (M), 208 g, and temperature change of water (N), 39 °C. The best way to report this value is 34,000 J, because the value for N, 39 °C, has only two significant figures. Thus according to the rule stated above, 33,941 J should be reported to two significant figures, or 34,000 J. (The uncertainty for 34,000 is located in the 1000s place, although the specific uncertainty is unknown for this process.) The proof is to work out the uncertainties, as above, and see that the last place in heat absorbed by water (O) and in the heat of combustion of candle wax (P) must be the thousands place, since that is itself uncertain. Students need to understand that to report the calculator answer of 33,941 J for heat absorbed by water (O) is a misrepresentation, since one cannot know it that precisely. To emphasize this, ask students how much the "missing" 59 J (34,000 J-33,941 J) would warm the more than 200 g water used. The approximate answer is only 0.07 °C, obviously unreadable on their thermometers.

Point out that uncertainty comes into experimental strategy all the time. In this case, for example, ask students why they were told to find the mass of the candle before and after burning to the nearest 0.01 g, while finding the mass of the tin can both empty and full of water only to the nearest 1 g. Why read the tin can only to ±1 g when, by taking a bit more time, these too could have been read to the nearest 0.01 g? The answer comes from uncertainties. Because loss of mass of the candle is small, even the ±0.02 g uncertainty is a large relative uncertainty. Using expected student results, it is 2%, while the can masses measured to ±1 g give a mass of water only 1% uncertain. Hence it makes no sense to spend time on those masses, since they already involve negligible uncertainties.

Teacher-Student Interaction

Move around the laboratory, initially pouring approximately 200-250 mL of the cold water (but no ice!) into each of the smaller tin cans and checking the arrangement of the apparatus. Be sure students quickly weigh the candle (± 0.01 g) and measure the water temperature. (The temperature initially should be approximately 4-5 °C.) Make sure students also immediately begin heating the cold water with the candle flame. As heating proceeds, remind students to stir gently, and to blow the candle flame out when the water temperature reaches approximately 45 °C. They must record the highest temperature reached by the water as its final temperature. Be sure students record all necessary data.

Anticipated Student Results

Raw Data	**Candle (± 0.01 g)**		**Tin can (± 1 g)**		**Temp. (± 0.5 °C)**	
	Before	*After*	*Empty*	*+ Water*	*Before*	*After*
	A	B	C	D	E	F
	45.97 g	45.06 g	61 g	269 g	6 °C	45 °C

Calculated results:

	Mass of candle burned (g)	**Mass of water warmed (g)**	**Temp. change of water (°C)**	**Heat absorbed by water (J)**	**Heat of combustion of candle wax (J/g)**
	L = A – B	M = D – C	N = F – E	O = M x N x 4.184 J/g°C	$P = \frac{O}{L}$
Value	0.91 g	208 g	39 °C	34,000 J	37,000 J
Uncertainty	± 0.02 g	± 2 g	± 1 °C	± 1200 J	± 2100 J
***% Uncertainty**	± 2 %	± 1 %	± 3 %	± 4 %	± 6 %

*More precisely, the % errors (E) are not additive but less since errors can be positive or negative: $E = \sqrt{E_1^2 + E_2^2}$

Answers to Data Analysis and Concept Development

1. See *Anticipated Student Results.*
2. See *Anticipated Student Results.*
3. As long as water is cooler than its room surroundings, heat will flow into water and warm it. When it is above room temperature, heat will leave the water and cool it. Since the temperature of the warmed water is as much above room temperature as the ice water was below room temperature, these errors cancel.
4. The uncertainties in the mass of candle burned (± 0.02 g), mass of water warmed (± 1 g) and temperature change are simply the sum of those in the data from which they are calculated. Uncertainties in heat absorbed by water and heat of combustion of candle wax are found from the sum of the percent uncertainties in mass of water warmed (± 1 %) and temperature change (± 3 %). Add the uncertainties (1% + 3%), and convert to uncertainty in heat absorbed by water by taking 4% of 34,000 J. (There are more precise computations for error analysis.)

5. All the heat given off by the candle is not measured. The can and surrounding air are also warmed by the candle, but not measured. This will result in an experimental heat of combustion that is too low. (Errors in measurement can result in random errors in the experimental heat of combustion.)

Answers to Implications and Applications

1. Table of information for other fuels:

Fuel	Heat of combustion (kJ/g)	Molar heat of combustion (kJ/mol)	Temperature rise for 1 L water per 1 g fuel (°C/L/g)
C	32.8	394	7.84
CH_4	55.7	891	13.3
C_2H_6	52.0	1560	12.4
C_3H_8	52.8	2330	12.6
C_2H_5OH	30.2	1390	7.22
$C_6H_{12}O_6$	15.5	2790	3.70
H_2	142	287	33.9
U-235	8.00×10^7	1.88×10^{10}	1.91×10^7

2. The relative value of uranium as a fuel compared to coal (mostly elemental carbon) is:

$$\frac{8.0 \times 10^7 \text{ kJ/g U-235}}{32.8 \text{ kJ/g C}} = 2.4 \times 10^6$$

times as much heat from U-235 as from an equal mass of coal.

3. ***Advantages***: Lightweight, compressible, plentiful (in seawater, if a source of energy to separate water into its elements is economical). Disadvantages: highly flammable (remember the Hindenberg!), "gas" tanks ruptured in collisions would result in cars becoming death traps

4. ***Advantages***: >2 Million times as much energy available on a mass basis as can be obtained from burning coal, much energy from little mass, no air pollution as in burning coal, "filler-up" at the tanks would be very infrequent. Disadvantages: radioactivity hazards, long-term handling, and storage of nuclear waste is very unclear presently

Post-Laboratory Discussion

Discuss the errors inherent in this activity. Get students involved in planning a strategy to minimize these errors. For example, ask them to justify starting with ice water that is to be warmed as far above room temperature as it was initially below that temperature. Why not start with room temperature and heat it to, say, 65 °C? (See Question 3 in *Answers to Data Analysis and Concept Development.*) An analogy may help here: Ask students what would happen in their house in winter if the furnace stops, or in summer if the air conditioner stops. (In the former, the heat would flow out of the house to the surroundings; in the latter, it would flow into the house from the outside.)

Ask if the heat measured (that entering the water) is all the heat given off by the candle. Point out that the can and surrounding air are also warmed by the candle, but not measured. Student values for the heat of combustion can therefore be expected to be too low. Indeed, the accepted value for paraffin wax is approximately 42,000 J/g.

Finally, be sure students understand the reason for dividing the number of joules released by the grams of candle burned. Point out that Student A might get 17,000 J by burning only 0.5 g candle, Student B might get 34,000 J by burning 1.0 g, Student C with 68,000 J from burning 2.0 g, yet each has the same heat of combustion: 34,000 J/g. This fact can be used to introduce ΔH units of kJ/mol later on in the module.

Possible Extensions

Ask students to design activities to measure heats of combustion of other fuels, such as sterno, rubbing alcohol and kerosene. *NOTE: For each of these materials, the mass loss during burning could be determined.* What about natural gas? How could results be standardized? The mass loss by the gas would be unavailable if it is taken from the laboratory gas jet, but the mass could be determined by collecting a sample over water. How about timing? Divide joules by seconds used for heating? Could these units be converted to joules per gram? How?

Ask students to research the procedure and apparatus used to obtain the "standard" values for heat of combustion located in handbooks, such as the *CRC Handbook of Chemistry and Physics.*

Assessing Laboratory Learning

Give students raw data from another fuel's heat of combustion experiment and ask them to design and complete a suitable "Calculated Results" table, ending with a heat of combustion as well as its uncertainty. Use data in *Answers to Implications and Applications* (Answer 1) to generate these "raw data."

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Heats of Reaction

Introduction

In this activity, you will observe and measure the heat of reaction for three different, but related, reactions. You will use your knowledge of the definition of heat to help you calculate the heat involved in each of these reactions. You will use the temperature change of a specified quantity of water contained in a glass Erlenmeyer flask, which will serve as the calorimeter. By knowing the mass of water and glass involved in the evolution or absorption of heat as well as the temperature change, you will be able to determine how much heat was exchanged in each of the reactions. These calculations will serve as a self-check on your experimental values.

To simplify the calculations somewhat, you will assume that the heat of reaction will raise the temperature of both the solution and the flask. We will neglect any heat lost to the surroundings. Recall that it takes 4.18 J to raise the temperature of 1 g water 1 °C. You may assume that the specific heat of solution is the same as that of pure water, 4.18 J/g °C. Glass has a specific heat of 0.84 J/g °C. The density of water is approximately 1.0 g/cm^3; thus the volume indirectly gives you the mass of water involved. You will use this information to calculate the heat of reaction.

Purpose

To determine the heat of reaction for several chemical reactions and then to use these values to calculate another ΔH.

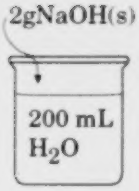
Safety

1. Wear protective goggles throughout the laboratory activity.
2. The sodium hydroxide pellets used in this activity are very caustic. It is imperative that you do not touch these pellets, as serious chemical burns could result. If sodium hydroxide is spilled on the hands or other area of the body, flush with running water for at least 10 min, or until the soapy feeling is gone. Notify your teacher immediately. Sodium hydroxide splashed into the eye can cause serious damage, even after irritation has stopped. Flush eyes in the eye wash fountain for at least 15 min. Have someone notify your teacher immediately.
3. The solutions of hydrochloric acid involved can be irritating to skin, eyes, and mucous membranes. If you spill any on yourself, immediately flush the area with water for several minutes and notify your teacher. If any is splashed in your eyes, flush your eyes with water immediately and continue doing this for at least 15 min. Use the eye wash fountain if one is available in the laboratory.
4. The use of a mercury thermometer adds the potential danger of a mercury spill. Mercury is toxic and very difficult to clean up. If someone should break a mercury thermometer, notify your teacher immediately so that the mercury can be cleaned up completely and disposed of properly.
5. The solutions at the end of the activity can be safely poured down the drain, and flushed with tap water.

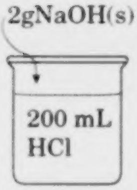
Procedure

Part I

1. Determine the mass of a 250-mL beaker, to the nearest 0.01 g.
2. Pour 200 mL (± 1 mL) of cool tap water into the beaker. Stir slowly and carefully with a thermometer until a constant temperature is reached (approximately room temperature). Record this temperature (± 0.2 °C).

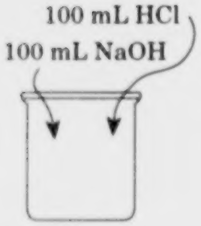


3. Determine the mass of about 2 g solid sodium hydroxide (NaOH), to the nearest 0.01 g. In order to do this as quickly as possible to prevent the solid from absorbing moisture from the air (and thus increasing its mass), your teacher will give you explicit instructions about how many pellets you will need to approximate 2 g (to stay within the range of 1.9 to 2.1 g).
4. Pour this measured sample of solid NaOH into the 200 mL of water already in the beaker. Stir the beaker vigorously (but avoid splashing) until the pellets are all dissolved. Measure the temperature and record the highest value observed.
5. Now dispose of this solution, and rinse the beaker in preparation for *Part II*.



Part II

1. Repeat Steps 1, 2, and 3 from *Part I*, except in place of the tap water used in Step 2, substitute 200 mL of 0.25 M hydrochloric acid (HCl) solution.
3. Now repeat Steps 4 and 5 from *Part I*, adding the measured NaOH to the 0.25 M HCl.



Part III

1. Measure 100 mL of 0.50 M HCl into a 250-mL beaker.
2. Measure 100 mL of 0.50 M NaOH into a second 250-mL beaker.
3. Measure the temperature of each of the solutions and record. (Rinse the thermometer before transferring it from one solution to the other.)
4. Pour the NaOH solution into the acid solution. Mix quickly and record the maximum temperature reached.
5. Thoroughly wash your hands before leaving the laboratory.

Data Tables

	Part I Reaction 1	*Part II* Reaction 2
Mass of beaker, g		
Mass of weighing paper, g		
Mass of paper + NaOH, g		
Initial temperature, °C		
Final temperature, °C		

	Part III Reaction 3
Initial temperature HCl soln, °C	
Initial temperature NaOH soln, °C	
Final temperature mixture, °C	

Data Analysis and Concept Development

Part I

Fill in the results table below with the information from the following calculations.

1. Calculate the change in temperature, ΔT, for the reaction.
2. Calculate the quantity of heat that is absorbed by the solution. (Remember that Q = m x C x ΔT.) Assume that the density of all solutions is 1.00 g/mL.
3. Calculate the quantity of heat absorbed by the beaker.
4. Calculate the total quantity of heat absorbed by the system.
5. Calculate the amount of NaOH used (in moles).
6. Calculate the total quantity of heat involved per mole of NaOH.
7. Express your results as the heat of reaction, ΔH, in kJ/mol.

Part II

Fill in the results table for your data in *Part II*. The calculations are the same as for *Part I*.

Part III

Fill in the results table for your data in *Part III*. The calculations are the same as for *Part I*.

Results Table

	Part I **Reaction 1**	***Part II*** **Reaction 2**	***Part III*** **Reaction 3**
1. Change in temperature, °C			
2. Heat absorbed by solution, J (200 g x 4.18 J/g °C x ΔT)			
3. Heat absorbed by beaker, J (Mass beaker x 0.84 J/g °C x ΔT)			
4. Total heat absorbed, J (Heat soln + heat beaker)			
5. Moles NaOH used in reaction			
6. Total heat per mole NaOH, kJ/mol NaOH			
7. Heat of reaction, kJ/mol NaOH			

Implications and Applications

1. Compare ΔH_2 with the sum of ΔH_1 and ΔH_3.
2. Calculate the percent difference between ΔH_2 and the sum of ΔH_1 and ΔH_3. (Assume ΔH_2 to be the correct value.)

3. Write the net ionic equations for Reactions 1, 2, and 3.
4. Each of the reactions above involves a specific type of heat of reaction. Reaction 1, for example, represents the heat of solution of solid NaOH. Write similar statements for the other two reactions.
5. Suppose you had used 4.0 g NaOH in Reaction 1, instead of 2.0 g.
   a. Calculate the quantity of heat involved.
   b. How would this result affect the ΔH, measured in kJ/mol, for this reaction?
6. How could we make use of the concept of heats of reaction in our everyday lives?

References

Similar activities are presented in *Chemistry: Experimental Foundations Lab Manual* (Merrill *et al.*, 1987, pp. 60-61) and *CHEMStudy Lab Manual* (1965, pp. 39-40).

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Heats of Reaction

Major Chemical Concept

This laboratory activity will provide data for students to calculate heats of reaction for three different reactions, and then to compare them in order to demonstrate Hess's Law.

Level

All college preparatory classes in chemistry.

Expected Student Background

Students should have already performed calorimetry calculations involving mass of water heated and its temperature change. The goal of this exercise is to take calorimetry one step farther and to compare different ΔH calculations to each other.

Time

All three parts of this activity can be completed in a 50-min period. Alternatively, specific parts could be assigned to individual laboratory groups, and their results could be combined to get a complete set of class data.

Safety

Read the *Safety Considerations* in the *Student Version*. You should be vigilant in observing student behavior throughout this series of activities. Be sure students have been shown the location of the eye wash facility prior to the start of the activity.

Materials (For 24 students working in pairs)

Nonconsumables (per laboratory team)

24 Beakers, 250-mL
12 Thermometers
12 Graduated cylinders, 100-mL
Centigram balance
12 Scoopulas
12 Stirring rods

Consumables

NOTE: Since this is a quantitative experiment, the solutions must be made up carefully to the proper concentrations. Ideally, the class solutions should be standardized.

Sodium hydroxide pellets, NaOH, 48 g
0.25 M Hydrochloric acid, HCl, 2 L (42 mL conc HCl diluted to 2 L)
0.5 M Hydrochloric acid, HCl, 1 L (42 mL conc HCl diluted to 1 L)
0.25 M Sodium hydroxide, NaOH, 1 L (10 g NaOH per 1 L solution)
0.5 M Sodium hydroxide, NaOH, 1 L (20 g NaOH per 1 L solution)
Tap water, 200 mL
Weighing paper

Advance Preparation

1. Solutions of HCl and NaOH must be prepared at least one day in advance so that they will be at room temperature on the day of the experiment. Both dilution reactions release heat, so the solutions will become hot when first prepared. Time must be allowed (at least overnight) for this heat to dissipate.

2. Since rather large volumes of acid and base solutions will be needed, you may want to employ siphons in large 4-L jugs with rubber tubing and pinch clamps to dispense the liquids to students.
3. The mass of the moisture picked up by the NaOH is apparently not significant, even though it appears that the pellets become quite wet with time. Results should not be affected by the slight (<1%) mass gain.
4. Ask students all to use the same labeled graduated cylinder (or two) for one solution, and keep it (them) at each jug, to keep dilution to a minimum and ensure that students are not contaminating their solutions *via* unwashed graduated cylinders. This method will require students to have only one cylinder at their laboratory station, to measure the water for *Part I*.

   Alternatively, you may want to have another siphon in a jug of slightly cool water that you have cooled ahead of time as a fourth solution. Assigning a labeled graduated cylinder or two at this fourth jug will negate the need for any cylinders at individual laboratory stations. Using this method, rinse each graduated cylinder first with a small volume of the assigned liquid and discard the liquid. From that point on, the only liquid that should be in that cylinder will be the correct (assigned) one.

Pre-Laboratory Discussion

1. Avoid mentioning Hess's Law to allow students to arrive at this "discovery" themselves.
2. Tell students the number of pellets required to approximate 2.0 g. Explain the method they are to use to clean up spilled pellets.
3. Warn students of the dangers of both hydrochloric acid and sodium hydroxide solutions, as well as solid sodium hydroxide.
4. Discuss with students the appropriate time for drying glassware (only when its mass needs to be determined). This means that the flask only needs to be dried once—the first time it is cleaned, because its mass will be determined immediately thereafter. Drying only when necessary will save students time.
5. Discuss with students the method you are using to dispense the solutions.

Teacher-Student Interaction

If assigned pre-laboratory preparations include making a data table, move around the room to ensure students have done this. (This is one sign that advanced preparations have been done by the student.) Also check to see that students understand the directions, especially any given verbally that deviate from the printed version.

Anticipated Student Results

	Part I **Reaction 1**	***Part II*** **Reaction 2**
Mass of beaker, g	125 g	125 g
Mass of weighing paper, g	0.89 g	0.88 g
Mass of paper + NaOH, g	2.97 g	2.80 g
Initial temperature, °C	24.0 °C	23.8 °C
Final temperature, °C	26.4 °C	29.6 °C

	Part III **Reaction 3**
Initial temperature HCl soln, °C	23.8 °C
Initial temperature NaOH soln, °C	24.2 °C
Final temperature mixture, °C	27.4 °C

Results Table

	Part I **Reaction 1**	***Part II*** **Reaction 2**	***Part III*** **Reaction 3**
1. Change in temperature, °C	2.4 °C	5.8 °C	3.4 °C
2. Heat absorbed by solution, J (200 g x 4.18 J/g °C x ΔT)	2.0 x 10^3 J	4.8 x 10^3 J	2.8 x 10^3 J
3. Heat absorbed by flask, J (Mass flask x 0.84 J/g °C x ΔT)	250 J	610 J	360 J
4. Total heat absorbed, J (Heat soln + heat beaker)	2.3 x 10^3 J	5.4 x 10^3 J	3.2 x 10^3 J
5. Moles NaOH used in reaction	0.052 mol	0.048 mol	0.050 mol
6. Total heat per mole NaOH, kJ/mol NaOH	44 kJ/mol	110 kJ/mol	64 kJ/mol
7. Heat of reaction, kJ/mol NaOH	–44 kJ/mol	–110 kJ/mol	–64 kJ/mol

Answers to Implications and Applications

1. ΔH_2 (–110 kJ/mol NaOH) is approximately equal to the sum of ΔH_1 and ΔH_3 (–108 kJ/mol NaOH).

2. $$\frac{[-110\ \text{kJ} - (-44\ \text{kJ} + -64\ \text{kJ})]}{-110\ \text{kJ}} \times 100\% = 2\%$$

   (This result is probably better than most students will get, but even 10-15% difference should be reasonable.)

3. $NaOH(s) + H^+(aq) \rightarrow H_2O(l) + Na^+(aq) + 110\ kJ$ (2)

   $OH^-(aq) + H^+(aq) \rightarrow H_2O(l) + 64\ kJ$ (3)

   (*NOTE: Values for ΔH will vary.*)

4. Reaction 2 represents the heat of reaction of solid NaOH and $H^+(aq)$ and is the sum of Reactions 1 and 3. Reaction 3 represents the heat of reaction for $H^+(aq)$ and $OH^-(aq)$, the heat of formation of water from its ions.

5. a. Using 4.0 g NaOH instead of 2.0 g would have produced twice as much heat, since twice as much reactant is involved.

   b. The ΔH, however, would not change, since twice the heat divided by twice the moles of reactant would still equal the same value.

6. One answer to this question was given in the *Content in a Nutshell*—hot packs and cold packs.

Post-Laboratory Discussion

You may need to review ΔH calculations with students. This may include the transition from kJ/mol to actual sign convention. The importance of Hess's Law should be stressed here; this regularity can be used to predict many heats of reaction. This activity should lead directly to an in-class discussion of Hess's Law (with problems to be solved), as well as a discussion of standard heats of formation and how students can use these to calculate heats of reaction.

Possible Extensions

You may want students to try other Hess's Law reactions, such as the heat of combustion of magnesium (*Laboratory Manual to Accompany Chemistry, Silver Burdett*, 1973, p. 56).

Assessing Laboratory Learning

Students can be given problems to calculate the heat of reaction for a specific reaction (several are listed below) from a table of standard heats of formation of compounds.

1. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = ? \qquad (1)$

Reactions for calculating the heat of reaction:

$C(s) + 2H_2(g) \rightarrow CH_4(g) \qquad \Delta H_f^o = -74.81\ kJ \qquad (2)$

$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H_f^o = -393.51\ kJ \qquad (3)$

$H_2(g) + {}^1/_2O_2(g) \rightarrow H_2O(l) \qquad \Delta H_f^o = -285.83\ kJ \qquad (4)$

$20C(s) + 21H_2(g) \rightarrow C_{20}H_{42}(s) \qquad \Delta H_f^o = -457.13\ kJ \qquad (5)$

2. $2C_{20}H_{42}(s) + 61O_2(g) \rightarrow 40CO_2(g) + 42H_2O(g) \qquad \Delta H_f^o = ?\ kJ$

(Use the same equations as in #1 above.)

3. $Na^+(g) + Cl^-(g) \rightarrow NaCl(s) \qquad H = ?kJ$

Reactions:

$Na(s) + {}^1/_2Cl_2(g) \rightarrow Na(g) + Cl(g) \qquad \Delta H_f^o = +230\ kJ \qquad (1)$

$Na(g) + Cl(g) \rightarrow Na^+(g) + Cl^-(g) \qquad \Delta H_f^o = +147\ kJ \qquad (2)$

$Na(s) + {}^1/_2Cl_2(g) \rightarrow NaCl(s) \qquad \Delta H_f^o = -411\ kJ \qquad (3)$

Answers

1. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = ? \qquad (1)$

$$\begin{aligned}
\Delta H &= \Delta H^o_{f\,products} - \Delta H^o_{f\,reactants} \\
&= [\Delta H^o_f CO_2 + 2\Delta H^o_f H_2O] - [\Delta H^o_f CH_4 + 2\Delta H^o_f O_2] \\
&= [1\ mol\ CO_2 \times \frac{-393.51\ kJ}{1\ mol\ CO_2} + 2\ mol\ H_2O \times \frac{-285.83\ kJ}{1\ mol\ H_2O}] - \\
&\qquad [1\ mol\ CH_4 \times \frac{-74.81\ kJ}{1\ mol\ CH_4} + 2\ mol\ O_2 \times \frac{0\ kJ}{1\ mol\ O_2}] \\
&= (-965.17\ kJ) - (-74.81\ kJ) \\
&= -890.36\ kJ
\end{aligned}$$

OR

$$\begin{array}{lll}
CH_4(g) \rightarrow C(s) + 2\,H_2(g) & \Delta H = +74.81\text{ kJ} & -(2) \\
C(s) + O_2(g) \rightarrow CO_2(g) & \Delta H = -393.51\text{ kJ} & (3) \\
2[H_2(g) + {}^1/_2O_2(g) \rightarrow H_2O(l)] & 2(\Delta H) = -571.66\text{ kJ} & 2 \times (4) \\
\hline
CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) & \Delta H = -890.36\text{ kJ} &
\end{array}$$

2. $2C_{20}H_{42}(s) + 61O_2(g) \rightarrow 40CO_2(g) + 42H_2O(l) \quad \Delta H = ?\text{ kJ}$

$$\begin{aligned}
\Delta H &= \Delta H^{0}_{f\,\text{products}} - \Delta H^{0}_{f\,\text{reactants}} \\
&= [40\Delta H^{0}_{f\,CO_2} + 42\Delta H^{0}_{f\,H_2O}] - [2\Delta H^{0}_{f\,C_{20}H_{42}} + 61\Delta H^{0}_{f\,O_2}] \\
&= [40\text{ mol }CO_2 \times \frac{-393.51\text{ kJ}}{1\text{ mol }CO_2} + 42\text{ mol }H_2O \times \frac{-285.83\text{ kJ}}{1\text{ mol }H_2O}] - \\
&\qquad [2\text{ mol }C_{20}H_{42} \times \frac{-457.13\text{ kJ}}{1\text{ mol }C_{20}H_{42}} + 61\text{ mol }O_2 \times \frac{0\text{ kJ}}{1\text{ mol }O_2}] \\
&= (-27745.26\text{ kJ}) - (-914.26\text{ kJ}) \\
&= -26831\text{ kJ}
\end{aligned}$$

OR

$$\begin{array}{lll}
2[C_{20}H_{42}(s) \rightarrow 20C(s) + 21H_2(g)] & \Delta H = +914.26\text{ kJ} & -2\ (5) \\
40[C(s) + O_2(g) \rightarrow CO_2(g)] & 40(\Delta H) = -15740.40\text{ kJ}40\ (3) & \\
42[H_2(g) + {}^1/_2O_2(g) \rightarrow H_2O(l) & 42(\Delta H) = -12004.86\text{ kJ}42\ (4) & \\
\hline
2C_{20}H_{42}(s) + 61O_2(g) \rightarrow 40CO_2(g) + 42H_2O(g) & \Delta H = -26831\text{ kJ} &
\end{array}$$

3.

$$\begin{array}{lll}
Na^+(g) + Cl^-(g) \rightarrow Na(g) + Cl(g) & \Delta H = -147\text{ kJ} & -(2) \\
Na(g) + Cl(g) \rightarrow Na(s) + {}^1/_2Cl_2(g) & \Delta H = -230\text{ kJ} & -(1) \\
Na(s) + {}^1/_2Cl_2(g) \rightarrow NaCl(s) & \Delta H = -411\text{ kJ} & (3) \\
\hline
Na^+(g) + Cl^-(g) \rightarrow NaCl(s) & \Delta H = -788\text{ kJ} &
\end{array}$$

DEMONSTRATIONS

With many of the following demonstrations, the teacher can use actual data from the tables of standard heats of formation to work with the class to calculate the heat of reaction involved in that specific demonstration. Always try all demonstrations ahead of time to ascertain the safest reaction conditions. Always wear safety goggles for chemical demonstrations.

Demonstration 1: Spontaneous Endothermic Reaction Involving Two Solids

Description

When two solids are mixed, an endothermic reaction occurs. Water in contact with the container freezes.

Materials

Ammonium chloride, NH_4Cl, 30 g
Barium hydroxide, $Ba(OH)_2 \cdot 8H_2O$, 30 g
Water
Erlenmeyer flask, 250-mL
Litmus paper
Wood block
Stirring rod

Safety

Ammonia gas is given off in the reaction. Students should be told not to inhale the gas but merely to waft some of the gas toward their noses with their hand to detect the smell.

Procedure

1. Place about 1 mL of water on the wood block.
2. Add approximately 30 g ammonium chloride and approximately 30 g barium hydroxide octahydrate to the Erlenmeyer flask.
3. Place the flask on the wetted wood block and stir vigorously.
4. After about 30 s, show students that the flask has frozen fast to the block of wood.

Remarks

The equation and the heat of reaction for the reaction above are:

$$2NH_4Cl(s) + Ba(OH)_2 \cdot 8H_2O(s) \rightarrow BaCl_2(aq) + 2NH_3(g) + 10H_2O(l) \ \Delta H = +\ 63.9 \text{ kJ}$$

Reference

This demonstration is described in: Summerlin and Ealy. (1988). *Chemical demonstrations, 1* (2nd Ed., p. 66). American Chemical Society.

Demonstration 2: Freezing Solid—Heats of Solvation

Description

A watch glass is "frozen" to a cork in an endothermic chemical reaction.

Materials

Sodium carbonate decahydrate, $Na_2CO_3 \cdot 10H_2O$ or ammonium chloride, NH_4Cl
Water
Cork
Watch glass

Safety

Use caution in allowing students to handle the system as it will be extremely cold. Such small amounts of solute are used that it is relatively safe to discard the solution down the drain.

Procedure

1. Place a watch glass onto a cork, upon which you previously placed several drops of water.
2. Dissolve some of one of the salts listed above in 1-2 mL of water placed in the watch glass.

Remarks

This demonstration is similar to the one above involving two solids, except that this is merely a solvation processes. As the endothermic solvation process proceeds, enough heat is lost by the water on the cork to freeze the water. The ice fuses the watch glass and cork together.

Demonstration 3: Hot and Cold Packs

Purpose

To illustrate that dissolution of solids can cause heat changes that are easily detected by feeling the container.

Materials

Calcium chloride, anhydrous, $CaCl_2$, 30 g
Ammonium chloride, NH_4Cl, 50 g
Water
2 Zip-closure bags
2 Vials with caps, 30-mL

Safety

The hot pack gets hot enough to cause discomfort. If you pass the cold and hot packs around the class, caution students not to hold the hot pack solution in their hand for too long. Contents can be flushed down the drain after the demonstration.

Procedure

Hot pack:

1. Measure approximately 30 g $CaCl_2$ into a sandwich-sized zip-closure bag.
2. Fill a small vial with about 25 mL of water and cap it.
3. Insert the sealed vial into the bag and seal the bag.
4. Keeping the zip-closure bag sealed, pop the cap off the vial from the outside. Pour the water into the solid.
5. Shake the bag to mix the contents.
6. Pass the system around the class and have students observe the temperature change.

Cold pack:

1. Repeat Steps 1-6 above using 50 g NH_4Cl (instead of $CaCl_2$).

Remarks

The heats of reaction for the two reactions above are:

$$NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq) \qquad H = +15.2\ kJ$$

$$CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) \qquad H = -82.8\ kJ$$

Relate these two reactions to the real-world use of hot and cold packs, *e.g.*, in sports injuries, hunting, *etc.*

Reference

Marsella, G. (1987, February). Hot and cold packs. *ChemMatters*, *5*(1), 7-12.

Demonstration 4: Gallon of Gas Reaction (Methane Gas Can Explosion)

Purpose

To safely demonstrate a potentially explosive reaction involving the rapid combustion of natural gas.

Materials

New, empty 1-gallon paint can
Masking tape, two 2" pieces
Rubber tubing
Methane or natural gas (mostly methane—propane won't work)
Matches
Long wax taper
Lucite explosion shield

Safety

Position the can so that the propelled lid will not impact with any student or breakable object (light fixtures, *etc.*). Use the explosion shield for protection (as described).

Procedure

1. Drill a 3/16" hole in the top of the lid of the paint can.
2. Drill a 1/2" hole about 1" above the bottom of the can.
3. Place the lid on the can and press it down tightly.

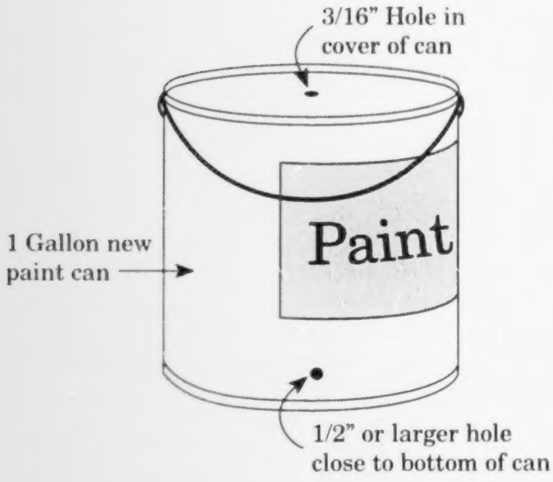

3/16" Hole in cover of can
1 Gallon new paint can
Paint
1/2" or larger hole close to bottom of can

***Figure 2. Diagram for* Demonstration 4.**

4. Attach one end of the rubber tubing to your laboratory gas jet or methane supply and the other end into the can (either hole). Since methane is lighter than air, the can should be above the point of inlet for methane.
5. Fill the can with gas and tape both holes shut with tape.
6. Place the can in an area that is not exposed to drafts, put the explosion shield in front of the can, and after removing both pieces of tape, light the gas at the top hole with the long taper.
7. Observe the size and color of the flame at the beginning, and compare with size and color as the reaction progresses. *NOTE: As the demonstration progresses, the flame becomes smaller and less luminous, even blue in color. Eventually it is just at the surface of the opening in the top of the can. At this point, any air current or movement of the can will cause the flame to extinguish. To prevent this, the can should be kept away from students and demonstrator alike, and surrounded, if possible, by a transparent plastic explosion shield to keep drafts of air away from the system.*
8. When the reaction is over, the can can be replenished immediately with gas in preparation for your next class. You can prepare the can in advance (even overnight), but be sure to tape the holes shut tightly.

Remarks

Even as the reaction proceeds, you can discuss the heat of reaction involved. This is a great example to illustrate Hess's Law of Heat Summation—students can see the results!

The reaction time is somewhat reproducible, typically about 10-12 min. If you time the demonstration as you practice, you will be able to predict rather closely the time of "extinguishing."

Equations involved are:

Overall reaction:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = ? \qquad (1)$$

Reactions for calculating the heat of reaction:

$$C(s) + 2H_2(g) \rightarrow CH_4(g) \qquad \Delta H^o_f = -74.81 \text{ kJ} \qquad (2)$$

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H^o_f = -393.51 \text{ kJ} \qquad (3)$$

$$H_2(g) + {}^1/_2\, O_2(g) \rightarrow 2H_2O(l) \qquad \Delta H^o_f = -285.83 \text{ kJ} \qquad (4)$$

Reverse Reaction 1 above, and add the three reactions to obtain the overall equation. Then add ΔH's to get –890.35 kJ for the overall reaction.

Answer:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = ? \qquad (1)$$

$$\begin{aligned}
\Delta H &= \Delta H^o_{f\,\text{products}} - H^o_{f\,\text{reactants}} \\
&= [\Delta H^o_f CO_2 + 2\Delta H^o_f H_2O] - [\Delta H^o_f CH_4 + 2\Delta H^o_f O_2] \\
&= [1 \text{ mol } CO_2 \times \frac{-393.51 \text{ kJ}}{1 \text{ mole } CO_2} + 2 \text{ mol } H_2O \times \frac{-285.83 \text{ kJ}}{1 \text{ mol } H_2O}] - \\
&\qquad [1 \text{ mol } CH_4 \times \frac{-74.81 \text{ kJ}}{1 \text{ mol } CH_4} + 2 \text{ mol } O_2 \times \frac{0 \text{ kJ}}{1 \text{ mol } O_2}] \\
&= (-965.17 \text{ kJ}) - (-74.81 \text{ kJ}) \\
&= -890.36 \text{ kJ}
\end{aligned}$$

Discussion can involve much more detail, including an enthalpy diagram (see *Transparency Masters 6* and *7* in the *Appendix*), and calculation of heat of reaction on the basis of bond dissociation energies.

Demonstration 5: Methanol—Tennis Ball Cannon

Purpose

To demonstrate the explosive nature of the reaction of methanol vapors, and to relate this to enthalpy and entropy changes.

Materials

Several cans of the same diameter, such as tennis ball cans or food cans of small diameter
1 Tennis ball (or Nerf ball)
Electrician's tape
Dropping bottle of methanol
Matches, long (if long matches are not available, use a wax taper)
Ringstand and support clamp

Safety

Be very careful to aim the "cannon" away from any person in the room. An explosion shield should be used in front of the taped stack of cans to prevent a rupture from propelling a can into the audience. Do not use plastic "cans"!

Procedure

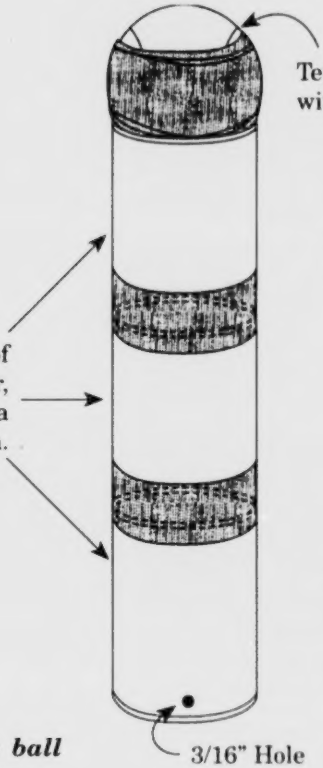

Tennis ball wrapped with electrical tape.
Three cans of identical diameter, taped to form a column 15-24" high.
3/16" Hole

Figure 3. Tennis ball can diagram.

1. Cut both the top and bottom from two small diameter cans, and one metal tennis ball can or two tennis ball cans.
2. Cut only one end off one can. (One end should be closed.)
3. Tape the (two or) three cans together with electrician's tape, going around each can several times to ensure a tight seal. Several cans of any height with a diameter approximately the same as that of a tennis ball can be taped together. (Apparently tennis ball cans are no longer made of metal, and the plastic varieties don't work very well.)
4. Drill a 5/16" hole about 1/2" from the bottom of the can with the one end remaining. Make sure the inside of the upper open can is filed smooth. (File the inside if necessary.)
5. Wrap the tennis ball with electrical tape until the ball fits snugly into the open end of the taped set of cans (the "cannon").
6. Place the tennis ball snugly into the opening of the stack of taped cans.
7. Squeeze two eyedroppers of methanol into the hole at the base of the bottom can.
8. Overturn the can and allow the methanol to evaporate. While you await the evaporation, mount the can with a clamp to a ringstand on the table. (You might want to tilt the cannon at an angle so that the ball's trajectory will not intercept anyone or any fragile object.)
9. Light a long match or wax taper and hold it at the hole in the bottom can. The methanol-air mixture will explode and drive the ball into the air. The effect is greatest when the tape-wrapped ball is tightly fitted into the top can.
10. An immediate repetition of the demonstration usually is a dud, illustrating that a new supply of air (oxygen) is necessary for the reaction to recur.

Remarks

At 298.2 K (25 °C):

$$\begin{array}{lrcll} 1. & 2CH_3OH(l) & \rightarrow & 2CH_3OH(g) & \Delta H = +37.5\ kJ \\ 2. & 2CH_3OH(g) + 3O_2(g) & \rightarrow & 2CO_2(g) + 4H_2O & \Delta H = -1343.1\ kJ \end{array}$$

Overall,

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O \qquad \Delta H = -1305.6\ kJ$$

The large increase in gas volume (resulting in the increased pressure that propels the ball) is correlated to an overall entropy change of +202.9 J/K.

Demonstration 6: Heat, Temperature, and Heat Capacity In Balloons

Purpose

To illustrate the large heat capacity of water as compared to air.

Materials

Balloons, two large, dark (black or violet)
Water
Candle
Two ringstands

Safety

Don't leave the flame under the second balloon for too long, or the balloon may develop a weak spot and burst. Discard balloons after use.

Procedure

1. Blow up one balloon to a diameter of 6-8 inches and tie a knot or seal it off by some other method.
2. Fill the second balloon part way with some water (approximately 50-100 mL will do) before blowing it up to the same approximate diameter as the first balloon. Tie this one off in the same way you did the first one.
3. Suspend the two balloons almost at the top of two separate ringstands. (If you have done all this correctly, unknowing students should not be able to detect the pool of water in the bottom of the second balloon.)
4. Ask students what they think will happen (prediction) when you hold a lighted candle under one balloon. Light a candle and test their hypothesis, holding the lit candle under the balloon filled only with air.
5. Now discuss (ostensibly) reproducibility of results and suggest the need to repeat the activity with the second balloon. Hold the lighted candle at the bottom of the second balloon where it will touch the center of the pool of water. Students will expect this balloon to burst also, but it will not.

Remarks

The second balloon will not burst because of the high heat capacity of the water. High heat capacity means that the water will absorb large quantities of heat (from the candle flame) without changing its temperature appreciably. The temperature of the burning candle is high (high enough to burst the material of the first balloon), whereas the temperature of the water in the second balloon is much lower (probably 20-25 °C). The temperature of the flame could be measured with a metal probe thermometer, while the water temperature can be easily measured by a standard laboratory thermometer.

Heat is that form of energy that flows across the boundary (synthetic rubber balloon) from the external high temperature to the lower temperature of the water. The water, rather than the balloon, absorbs the heat.

A similar activity can be done using a paper cup as the container in which to boil water. If one maintains a low flame, even on a burner, the water will boil with little or no charring of the paper. Be sure you use a cup that is not coated with wax or plastic to avoid the wax or plastic burning first!

Demonstration 7: Burning Candles—Reaction Mechanisms and Concurrent Reactions—Oxidation

Purpose

To develop, on the submicroscopic level, an understanding of what is occurring when a candle burns.

Materials

Several candles
One candle with exposed wick at both ends
Florence flask (or round bottom flask) containing ice
Wide-mouth Erlenmeyer flask
Two support clamps
Matches
Ice

Safety

As always, be careful when using flammable substances. Be especially cautious when burning your candle at both ends!

Procedure

Part 1

1. Light a candle and allow it to form a good-sized pool of wax.
2. Light another match, blow out the candle and then hold the lighted match above the wick, not touching the wick. The candle should relight without touching the match.

Part 2

1. Hold the round flask containing the ice a few inches above a lighted candle. A black deposit of soot (unburned carbon) will start forming, and perhaps moisture will condense.

Part 3

1. Place a clamp on the neck of the Erlenmeyer flask and hold the upended flask so that the neck opening is above the burning candle. Moisture will collect on the **inside** walls of the flask.

Part 4

1. Light one end of the double-wick candle holding it with a clamp, while concealing the other wick from the class.
2. Ask them, "How could you measure the rate at which a candle burns?" *[Measure height change, measure mass decrease,* etc.*]*
3. Then ask them, "How would you increase the rate at which the candle burns?" *[Use pure oxygen gas.]*
4. Then light the candle at the other end to show yet another way to increase the rate. A quote from Dr. Hubert N. Alyea, Professor Emeritus of Princeton University, fits here: "If you burn your candle at both ends, you will only live half as long . . . but you'll have more than twice as much fun!"

Remarks

The concurrent thermochemical reactions are:

$C_{20}H_{42}(s) \rightarrow C_{20}H_{42}(l) \rightarrow C_{20}H_{42}(g)$ (1)

$2C_{20}H_{42}(s) + 21O_2(g) \rightarrow 40C(s) + 42H_2O(g) \quad \Delta H = -11091\ kJ$ (2)

$2C_{20}H_{42}(s) + 41O_2(g) \rightarrow 40CO(g) + 42H_2O(g) \quad \Delta H = -15512\ kJ$ (3)

$2C_{20}H_{42}(s) + 61O_2(g) \rightarrow 40CO_2(g) + 42H_2O(g) \quad \Delta H = -26831\ kJ$ (4)

Part 1 of the procedure can be explained by observing Equations 1, 2, and 3 above, all of which contain unburned substances [$C_{20}H_{42}(g)$, C(s), and CO(g)] that are available for further burning if the energy of activation is provided (as by the match).

Part 2 of the demonstration shows Reaction 2 above, where unburned carbon deposits on the flask bottom.

Part 3 of the demonstration shows the production of water vapor that condenses inside the flask. This is shown in Reactions 2, 3, and 4.

Part 4 of the demonstration involves the addition of oxygen to the reaction system. Oxygen is the limiting reagent in the burning of a candle, as Reactions 2 and 3 show. Only when an abundance of oxygen is present does the paraffin burn completely to carbon dioxide and water vapor. Typically, Reaction 2 above is the major combustion reaction for candles, as evidenced by the yellow flame (incandescent carbon particles) and the evolution of smoke or soot (unburned carbon).

Questions you might want to ask students:

1. What are the functions of the wick? *[To light the candle, to melt the wax, to draw up molten wax by capillary action, to vaporize the wax.]*
2. How do we know that wax vapor is undergoing combustion? *[Remember the match above the just-burned wick?]*
3. Which is the major combustion reaction in the candle? *[Reaction 2]*
4. Why is this the major combustion reaction? *[Oxygen gas is the limiting reagent.]*
5. Discuss the mechanism of burning of the candle up to Reaction 2 above. *[paraffin(s) → paraffin(l) → paraffin(g)]*
6. What makes the mechanisms of Reactions 2, 3, and 4 difficult to determine? *[The sheer number of molecules involved in the reaction coupled with the complexity of the bond breaking and bond forming steps on a molecular level.]*
7. Which are consecutive phase changes? *[The two reactions included in Reaction 1 illustrate consecutive reactions.]*
8. Which are concurrent reactions? *[Reactions 2, 3, and 4 illustrate concurrent reactions.]*

Two questions that could be cross-referenced in *Links and Connections* section:

1. Contrast Reactions 2, 3, and 4 in terms of energy production and air pollution. [*Obviously, Reaction 4 produces the most energy (26831 kJ* vs. *11091 kJ and 15512 kJ), and it is also the least polluting of the three reactions, although carbon dioxide is now known to contribute to the "greenhouse effect." The other two reactions each produce less energy per mole of paraffin burned, and they contaminate the air by producing dirty soot and carbon monoxide.]*

2. In the burning of natural gas (consisting of CH_4, C_2H_6, C_3H_8, and C_4H_{10}) in a burner, gas stove, gas oven or gas heater, the blue flame indicates a much more energy-efficient burning process. Why is the efficiency of this process so much greater than that of the combustion of candle wax? Use the equations for the combustion reaction for paraffin to help explain this. Think mechanisms. *[The main reason is the premixing of natural gas with air (oxygen). Gasoline (e.g., C_8H_{18}) burns more efficiently because it is premixed in the carburetor. The four gases involved in natural gas contain far fewer carbon and hydrogen atoms than paraffin, and thus occur in much less complex reaction mechanisms than paraffin. Also, since they are already gases, whereas paraffin is solid, far fewer complicated mechanisms are needed for bond breaking and reforming with natural gas than with paraffin.]*

References

See *Demonstration 1* in the *Reaction Rates* module for other burning reactions that have thermochemistry applications.

ΔH's for several of those reactions follow here:

Burning of steel wool:

$$3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s) \qquad \Delta H = -1118 \text{ kJ}$$

Burning of magnesium powder:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s) \qquad \Delta H = -1204 \text{ kJ}$$

These reactions with additional data could be used to illustrate the Born-Haber cycle. Since gases are consumed in both reactions, the entropy changes are negative and are unfavorable driving forces for the reaction. The latter reaction is employed in flash bulbs and cubes. Boy Scouts are taught to light fires with the steel wool/iron reaction. The lycopodium powder reaction has been used by magicians under the name, "dragon's breath." These reactions can be also used to introduce balancing of oxidation-reduction reactions.

Other Demonstration Ideas

1. **Explosive Reaction**

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) + \text{energy!}$$

See: Summerlin and Ealy. (1988). *Chemical demonstrations* (p. 27). American Chemical Society.

2. **Electrolysis of Water in Hoffman Apparatus**

Do this demonstration in concert with the previous demonstration above, emphasizing the large quantity of electrical energy needed to produce (and thus stored in) the small amounts of hydrogen and oxygen gases involved in the demonstration (see *Demonstration 4* in *The Mole* module.) You might also want to explode bubbles of a mixture of hydrogen and oxygen to show that large quantities of energy are released when these two elements recombine to form water.

Hubert Alyea, Professor Emeritus in Chemistry from Princeton, a master of demonstrations, did a demonstration involving a hydrogen/oxygen "pop" bottle. The demonstration consisted of a stoichiometric ratio (2:1) of hydrogen and oxygen in a stoppered soda "pop" bottle. He lit the mixture (aim away from the audience), then explained that this is the reason it is called a "pop" bottle! He also showed that water was produced in the reaction by pouring out several milliliters of water after the reaction. (He had poured the water in behind the scenes before the demonstration!) Much discussion has ensued recently concerning the safety of the demonstration and the probability of the bottle's exploding. No conclusion has been reached, but you would be wise to tape the entire bottle in clear tape or netting to minimize the risk. A plastic pop bottle might be a wise choice.

3. **Spontaneous Exothermic Reactions**

   See *Chemical Demonstrations 41* "Exothermic Reaction: Calcium Chloride" (Summerlin & Ealy, 1988, p. 67) and *Chemical Demonstration 42* "Exothermic Reaction: Sodium Sulfite and Bleach" (Summerlin & Ealy, 1988, p. 68).

4. **Entropy Increase during Endothermic Reaction**

   See Demonstration 26 in *Chemistry: Experimental Foundations, Teachers' Guide* (Parry *et al.*, 1987, pp. 381-382).

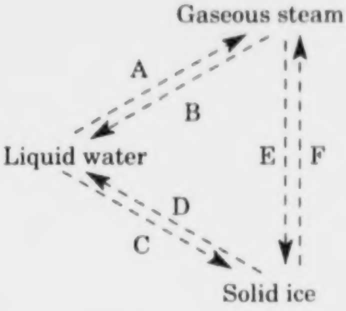
GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What is the meaning of the terms "surroundings" and "system"? *[The system is the portion of the universe under investigation. The surroundings are the rest of the universe.]*
2. What is the difference between heat and work? *[Heat is a form of energy that crosses a boundary whenever there is a difference in temperature at the boundary. The rate at which heat energy is transferred depends upon many factors including the thermal conductivity of the barrier and the temperature differential, ΔT. Work is any quantity of energy that flows across the boundary of a system during a change in state and is completely convertible into the lifting of a mass in the surroundings. $w = mgh$ (where w = work, m = mass, g = acceleration due to gravity, and h = height the mass was lifted); if the mass is lifted, h is positive and we say work has flowed into the surroundings.]*
3. What is the difference between temperature, heat, and heat capacity? *[See Misconception 1.]*
4. What is the difference between an endothermic reaction and an exothermic reaction? Give at least one example of each. *[In endothermic reactions heat is added to the system, whereas in exothermic reactions the system loses thermal energy. Examples of these two types of reactions are given throughout this module.]*
5. Consider the bonds, C–C, C=C, and C≡C. Which bond requires the greatest amount of energy to break? Which bond is the shortest? *[The triple bond, C≡C, is the strongest because, in essence, there are three bonds to be broken here. For the same reason, the triple bond is the shortest bond, since the forces of attraction between these two carbon atoms is the strongest.]*

Data	*Bond dissociation energy*	*Bond length*
C–C	*348 kJ/mol*	*0.154 nm*
C=C	*614 kJ/mol*	*0.133 nm*
C≡C	*812 kJ/mol*	*0.120 nm*

6. Examine the diagram:



   a. What names are given to the changes A, B, C, D, E, and F? *[A = evaporation, B = condensation, C = freezing, D = fusion (melting), E = deposition (desublimation?), F = sublimation]*
   b. Which changes are exothermic? *[B, C, E]*
   c. Which changes are endothermic? *[A, D, F]*
   d. Define the terms molar heat of fusion and molar heat of vaporization. *[Molar heat of fusion is the heat required to convert one mole of solid to liquid. Molar heat of vaporization is the heat required to convert one mole of liquid to gas.]*

7. Which would have the higher entropy?
   a. $H_2O(l)$ or $H_2O(g)$? *[$H_2O(g)$]*
   b. $Cl_2(g)$ or $2Cl(g)$? *[$2Cl(g)$]*
   c. $CO_2(s)$ or $CO_2(g)$? *[$CO_2(g)$]*
   d. $Ag^+(g)$ or $Ag^+(aq)$? *[$Ag^+(g)$]*

8. Which would have the higher energy?
   a. $H_2O(l)$ or $H_2O(g)$? *[$H_2O(g)$]*
   b. $Cl_2(g)$ or $2Cl(g)$? *[$2Cl(g)$]*
   c. $CO_2(s)$ or $CO_2(g)$? *[$CO_2(g)$]*
   d. $Ag^+(g)$ or $Ag^+(aq)$? *[$Ag^+(g)$]*

9. Write an equation for the complete combustion of:
   a. **Methane** $CH_4(g)$ *[$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$]*
   b. **Propane** $C_3H_8(g)$ *[$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$]*
   c. **Butane** $C_4H_{10}(g)$ *[$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$]*
   d. **Octane** $C_8H_{18}(l)$ *[$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$]*

10. It takes 571.5 kJ of energy to decompose two moles of water. How much energy is needed to decompose 1000 g water?

$$1000\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.0\ g\ H_2O} \times \frac{571.5\ kJ}{2\ mol\ H_2O} = 15875\ kJ \cong 2 \times 10^4\ kJ$$

Counterintuitive Examples

1. A spontaneous endothermic reaction seems to be in opposition to the tendency to minimum energy. It can be explained using the tendency toward maximum entropy. If the quantity, $T\Delta S$ is positive and greater in magnitude than ΔH, the reaction can be spontaneous ($\Delta G < 0$) despite its endothermicity. ($\Delta G = \Delta H - T\Delta S$)
2. Household dehumidifiers do not cool a home, yet a dehumidified home feels cooler than a nondehumidified home, even though both homes are at the same temperature because evaporation of water from the skin carrying away heat from the body occurs more rapidly in a low humidity environment.

Metaphors and Analogies

1. Use the analogy of a book on the floor *vs.* the same book on an upper shelf of a bookcase to demonstrate the energy of a substance changing as it reacts or changes.
2. See Item 1 in *Common Student Misconceptions.* Refer to the mechanical analogy for endothermic and exothermic reactions detailed in *Metaphors and Analogies* in the *Rates of Reaction* module.

TIPS FOR THE TEACHER

Language of Chemistry

adiabatic process process in which no heat exchange occurs between the system and its surroundings.

bomb calorimeter device used to measure the heat of a combustion reaction. The quantity measured is the heat of reaction at constant volume, $q_v = \Delta E$.

bond energy quantity of energy required to break a particular bond in one mole of gaseous molecules.

calorimeter device used to measure the quantity of heat exchanged between a system and its surroundings.

calorimetry measurement of heat changes.

chemical energy energy stored within the structural units of chemical substances.

closed system system that allows the exchange of energy (usually in the form of heat) but not mass with its surroundings.

endothermic process process that absorbs heat from its surroundings

energy capacity to do work or to produce change.

enthalpy (H) thermodynamic quantity used to describe heat content of a system at constant pressure.

enthalpy change difference in enthalpy between two states of a system at constant pressure.

enthalpy of formation enthalpy change accompanying the formation of a mole of a compound from the most stable forms of its elements.

enthalpy of reaction difference between the enthalpies of the products and the enthalpies of the reactants.

enthalpy of solution heat generated or absorbed when a certain amount of solute is dissolved in a certain amount of solvent.

entropy (S) direct measure of the randomness or disorder of a system.

entropy change extent to which the degree of order changes as the result of some process (Positive ΔS means an increase in disorder of the system.

entropy change of universe total entropy change, both system and surroundings, for a given process. For every spontaneous change, $\Delta S_{univ} > 0$.

exothermic process process that gives off heat to the surroundings.

first law of thermodynamics energy can be converted from one form to another, but cannot be created or destroyed

free energy (G) energy available to do useful work.

heat form of energy transfer between two bodies at different temperatures.

heat capacity quantity of heat required to change the temperature of an object or a given quantity of a substance by one degree Celsius.

heat engine device for converting heat into work. The engine absorbs heat at a high temperature, converts part of it to work, and discharges the remaining heat to the surroundings at a lower temperature.

heat of dilution heat change (absorption or evolution) associated with the dilution process.

heat of hydration heat change (absorption or evolution) associated with the hydration process.

heat of reaction quantity of heat associated with a chemical reaction.

heat of solution (See enthalpy of solution)

Hess's law enthalpy change for an overall or net process is the sum of enthalpy changes for individual steps in the process.

internal energy total energy attributed to the particles of matter and their interactions within a system.

joule (J) basic SI unit of energy, is the quantity of work done when a force of one newton acts through a distance of one meter.

kinetic energy energy available due to the motion of an object.

molar heat of fusion energy, expressed in kJ, required to melt one mole of a solid.

molar heat of sublimation energy (in kJ) required to sublime one mole of a solid.

molar heat of vaporization energy (in kJ) required to vaporize one mole of liquid.

phase change transformation from one phase to another.

potential energy energy available due to the position of an object.

second law of thermodynamics entropy of the universe increases in a spontaneous process and remains unchanged in an equilibrium process.

specific heat quantity of heat required to change the temperature of one gram of substance by one degree Celsius.

state of a system the values of all pertinent macroscopic variables of a system, *e.g.*, volume, pressure, temperature and composition.

surroundings that portion of the universe with which a system interacts.

system that portion of the universe selected for a thermodynamic study.

temperature measure of the average molecular kinetic energy of a substance.

thermal energy energy associated with the random motion of atoms and molecules.

thermochemical equation equation that shows both mass and enthalpy relations.

thermochemistry the study of the interconversion of heat and other forms of energy in chemical systems.

third law of thermodynamics entropy of a pure perfect crystal is zero at the absolute zero for temperature, 0 K.

work form of energy transfer that can be expressed as a force acting through a distance.

Common Student Misconceptions

1. **"Temperature and heat are the same."**

   Students commonly confuse the terms temperature and heat. Temperature is a measure of molecular motion. Heat is a form of energy that flows across a conducting boundary from the higher to the lower temperature. Using units, the difference in these terms is apparent:

   - Temperature—°C or K
   - Heat—J or kJ (or cal or kcal)

   Another example is to have them consider two containers of water, a teacup and a bathtub. Both containers are at the same temperature. Ask students which will cool down fastest. (Obviously, the teacup will be cooled down most quickly, due to the greater amount of heat contained in the bathtub of water relative to that of the teacup.)

   Yet another example is to note the similarity and the difference between a burning match and a bonfire. The two temperatures are relatively the same, but the heat content of each is vastly different.

2. **"Heat capacity, specific heat capacity and molar heat capacity are all the same term."**

   These three terms have similar, but different meanings. **Heat capacity** is the amount of heat needed to raise the temperature of a given amount of substance one degree Celsius (or one Kelvin). If 1 gram of the substance is used, the appropriate term for the heat involved is the **specific heat capacity**. For one mole of the substance, the heat involved is the **molar heat capacity**. Unit analysis will help to show the differences:

   - Heat capacity—J/C° or J/K
   - Specific heat capacity—J/g-C° or J/g-K
   - Molar heat capacity—kJ/mol-K or J/mol-K

   Two examples that will help students to understand these differences are the following:

   a. It takes 50.2 J to raise the temperature of 1.0 g of water 12.0 K (1 g x 4.184 J/g-K x 12.0 K); it takes 2510 J to raise the temperature of 50.0 g of water 12.0 K. It takes 904 J to raise the temperature of 1 mole (18 g) of water 12.0 K. Contrast the temperature change, T, and the heat input, q, for each change.

   b. It takes only 4.78 J to raise the temperature of 1.0 g of copper by 12.0 K; it takes 239 J to raise the temperature of 50.0 g copper by 12.0 K (50.0 g x 0.39 J/g-K x 12.0 K); and it takes 297 J to raise the temperature of 1 mole of copper (63.5 g) 12.0 K. Contrast the temperature change, T, and the heat input, q, for each change.

   c. Compare and contrast "a" and "b" above.

3. **"A mechanical engine can be made 100% efficient." (Automakers could make 100% efficient engines if they wanted to.)**

   The % efficiency is given by:
   $$\%E = \frac{T_{\text{hot engine}} - T_{\text{exhaust}}}{T_{\text{hot engine}}} \times 100$$

   and, in addition, there are always losses due to friction and other factors. Thus, a 100% efficient engine is impossible since there is always a temperature difference between the engine and the exhaust.

4. **"Heat energy can be converted completely into work."**

   Since work can be converted completely into heat, as shown by James Joule and Julius Mayer around 1843, it seems reasonable to assume that heat energy can be converted completely into work, since energy is conserved. This, of course, is **not** the case (No. 3 is an example).

5. **"If a reaction mixture gets hot, the reaction must be endothermic."**

   Apparently the reasoning is as follows: The mixture is hot; therefore, it must have absorbed heat. That makes it an endothermic reaction. Actually, exactly the opposite is true. The reaction is generating heat, and the reaction is exothermic, *i.e.*, heat flows out of the system to heat up the surroundings. This is a very pervasive misconception. Be careful to distinguish between system (loses energy) and surroundings (gains energy and becomes hot).

6. **"The entropy of an element is 0 at 298.2 K and 1 atmosphere."**

   The enthalpy and free energy of the elements at 298.2 K and 1 atmosphere (1 bar) are both 0, but the entropy, an absolute quantity, is not 0 at these conditions. (According to the third law of thermodynamics, the entropy of pure crystalline substances is 0 at 0 K.)

7. **"A calorie must be a huge quantity of energy—look how long it takes to burn off 1 calorie when I exercise."**

   The calorie one burns when exercising is a nutritionist's calorie, actually 1 kilocalorie or 1000 of a chemist's calories. The symbol for the food calorie is C (large "C"), representing 1000 calories (small "c"). Remember the calorie was originally defined as the quantity of heat needed to raise the temperature of 1 gram of water 1 C°.

Problem Solving

1. Calculate the heat of reaction, ΔH, for the neutralization of aqueous sulfuric acid by aqueous sodium hydroxide:

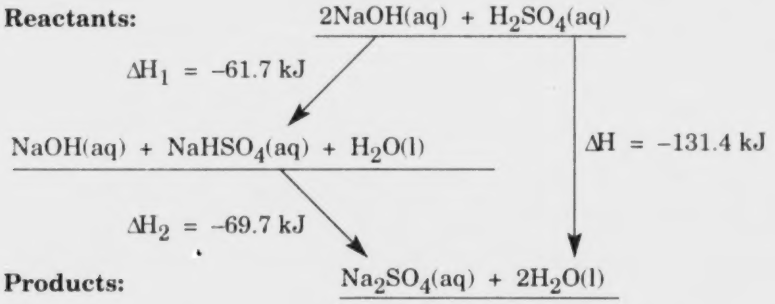
   $$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O \qquad \Delta H = ?$$

   From:

   $$NaOH(aq) + H_2SO_4(aq) \rightarrow NaHSO_4(aq) + H_2O \qquad \Delta H = -61.7\ kJ$$
   $$NaHSO_4(aq) + NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O \qquad \Delta H = -69.7\ kJ$$

   [$\Delta H = -131.4$ kJ]

2. Draw an enthalpy diagram to graphically illustrate the heats of the reaction in Question 1.



3. The fuel used for laboratory burners may contain mostly methane, $CH_4(g)$. Calculate the molar heat of combustion of methane, using Hess's Law, from the following data. The equation for the complete combustion of methane is:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) \qquad \Delta H = ?$$

Contributing heat of formation reactions:

a. $C(s) + 2H_2(g) \rightarrow CH_4(g) \qquad \Delta H = -74.8\ \text{kJ}$

b. $C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -393.5\ \text{kJ}$

c. $H_2(g) + {}^1/_2O_2(g) \rightarrow H_2O(g) \qquad \Delta H = -241.8\ \text{kJ}$

[First, reverse Reaction a; next, multiply Reaction c by 2 (for two moles of water formed); then add a, b, and c to get the overall equation, adding ΔH's to obtain –802.3 kJ for the overall reaction.]

4. Tungsten carbide, WC, is an extremely hard substance commonly found on the cutting surfaces of home and industrial saws and other cutting blades. Calculate the standard heat of formation of tungsten carbide, WC, from its elements, tungsten and graphite (carbon), using the data given below. The equation for the formation of tungsten carbide from its elements is:

$$W(s) + C(s) \rightarrow WC(s) \qquad \Delta H = ?$$

	ΔH, kJ
$C(s) + O_2(g) \rightarrow CO_2(g)$	–393.5
$2WC(s) + 5O_2(g) \rightarrow 2WO_3(s) + 2CO_2(g)$	–2391.6
$W(s) + 3O_2(g) \rightarrow 2WO_3(s)$	–1674.9

[–35.1 kJ/mol WC]

5. Many cigarette lighters contain liquid butane that, when burned, forms $CO_2(g)$ and $H_2O(g)$ according to the equation:

$$2C_4H_{10}(l) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$$

Given the following molar heats of formation at standard conditions, calculate:

a. the heat of combustion of two moles of butane.

b. the heat of combustion of one mole of butane.

c the heat of combustion of 1.0 g butane.

d. the heat of reaction for two moles of butane if the product were $CO(g)$ instead of $CO_2(g)$.

	ΔH^0_f
$C_4H_{10}(l)$	–127.4 kJ
$CO(g)$	–110.5 kJ
$CO_2(g)$	–393.5 kJ
$H_2O(g)$	–241.8 kJ
$H_2O(l)$	–285.8 kJ

[a. –5311.2 kJ; b. –2655.6 kJ; c. –45.7 kJ; d. –3047.2 kJ]

6. When 0.150 g of liquid octane, $C_8H_{18}(l)$ (molar mass = 114), was burned in a calorimeter containing 1500 g (weighed to the nearest gram) of water, the temperature rose from 24.146 °C to 25.286 °C. Calculate the heat of combustion (kJ/mol) of octane at 25 °C. *[The specific heat capacity of water is 4.184 J* $g^{-1}{}^\circ C^{-1}$*.]*

$$Q = m \times C \times \Delta T$$

$$1500\ \text{g} \times 4.184\ \text{J g}^{-1}{}^\circ\text{C}^{-1} \times 1.140\ {}^\circ\text{C}$$

$$Q = 7155\ \text{J}$$

and

$$0.150\ \text{g} \times \frac{1\ \text{mol } C_8H_{18}}{114\ \text{g } C_8H_{18}} = 0.00132\ \text{mol } C_8H_{18}$$

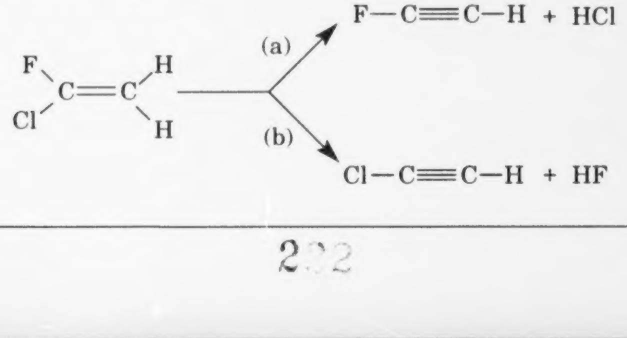
$$\Delta H = -Q_p/\text{mol } C_8H_{18} = \frac{-7155\ \text{J}}{0.00132\ \text{mol } C_8H_{18}} \times \frac{1\ \text{kJ}}{1000\ \text{J}}$$

$$\Delta H = -5420\ \text{kJ/mol } C_8H_{18}$$

7. Kinetic energy may be translational (T), rotational (R), or vibrational (V). Indicate what type(s) of kinetic energy is (are) involved in each of the following:
   a. Nolan Ryan throws a curve ball. *[T, R]*
   b. Nadja Salerno-Sonnenberg plucks a violin string. *[V for string; T for finger]*
   c. James ("007") Bond plays the bongo drums during a sky diving descent. *[T, V, R]*
   d. The earth moves around the sun. *[T, R]*
   e. A stainless steel ball traverses the deck of a pinball machine. *[T, R]*
   f. A ball attached to a spring is placed in a spinning orbital motion as it hangs from the end of an overhead steel arm. *[T, V, R]*
   g. Steffi Graf executes a backhand at Wimbledon. *[T, V, R]*

8. Indicate whether the following processes are exothermic (EX) or endothermic (EN) or no heat effect (N).
   a. A tree grows in the forest. *[EN]*
   b. A tree decays in the forest. *[EX]*
   c. A tree burns in the forest. *[EX]*
   d. Water is formed from the liquid oxygen-liquid hydrogen reaction in the engine of the Discovery rocket. *[EX]*
   e. The water droplets in the rocket exhaust plume evaporate. *[EN]*
   f. The water droplets in the rocket plume freeze. *[EX]*
   g. The water droplets in the rocket plume freeze and then remelt. *[N]*
   h. A cupful of sugar falls to the floor. *[EX]*

9. A laser may cause either of the following two chemical changes:



Calculate the heat of each reaction, ΔH, with the aid of the following average bond dissociation energies (kJ):

Bond	Bond Dissociation Energy (kJ/mol)
C–F	485
C–Cl	339
C–H	413
H–F	565
C=C	607
C≡C	835
C–C	345
H–Cl	431

[Recall that + sign denotes energy added to break a bond while – sign denotes energy given off as a bond forms:

a. + C=C + 607 kJ
+ C–Cl + 339 kJ   –C≡C –835 kJ
+ C–H + 413 kJ   –H–Cl –431 kJ
+ 1359 kJ   – 1266 kJ   net ΔH = + 93 kJ

b. + C=C + 607 kJ
+ C–F + 485 kJ   –C≡C –835 kJ
+ C–H + 413 kJ   –H–F –565 kJ
+ 1505 kJ   –1400 kJ   net ΔH = + 105 kJ]

10. A typical hamburger has an energy value of 15.0 kJ/g. If a student ate two half-pound hamburgers for lunch (1 lb = 454 g), and if none of the energy was stored in the student's body, how much water must the student sweat off in perspiration so that his body temperature remains constant? The heat of vaporization of water is 2260 J/g. *[3.01 kg water]*

11. Entropy is related to the number of ways of arranging a system in space, a measure of the degree of disorder. Predict whether the entropy will increase (I) or decrease (D) for each of the following changes:
    a. Air is pumped into a tire. *[D]*
    b. Gasoline vaporizes in the carburetor of a car. *[I]*
    c. Sugar is dissolved in a cup of coffee. *[I]*
    d. Moisture condenses on the windows of a car. *[D]*
    e. A new deck of playing cards is shuffled. *[I]*
    f. A cylinder of helium is used to fill 3000 balloons for the Thanksgiving Day Parade in New York City. *[I]*
    g. Snowflakes form in the upper troposphere. *[D]*
    h. Magnesium combines with oxygen gas to form magnesium oxide. *[D]*
    i. A box of black jelly beans is mixed with a box of green jelly beans. *[I]*

Decision Making

1. Each person, family, community, nation, and the entire world, must develop a program that maximizes energy conservation and minimizes environmental damage. Our present use of fossil fuels is both massive and wasteful. Our future petrochemical industries must have this precious commodity. Scientific, economic, political, and social considerations must be brought to bear on the problem. Available options are suggested below. You may wish to add others. Examine each option in terms of energy needs based upon statistical data and the best up-to-date information and recommend options that strike a proper balance between energy consumption and the reduction of environmental damage. You may wish to restrict your study to your specific locality.
   a. Accelerate research on nuclear fusion and seek new nuclear power technology.
   b. Explore increasing the production of synthetic fuels (synfuels) from coal. (Coal gasification and liquefication are some possible examples.)
   c. Evaluate the further development of oil from oil shale and tar sands.
   d. Based on new technology and increased safety factors resulting from European advances in nuclear technology, consider new nuclear fission power plants.
   e. Improve methodology and engine design that decreases fuel consumption in motor vehicles and aircraft.
   f. Reconsider the use of animals to replace machines for performing work. (Horses pulling wagons is an example.)
   g. Set up a public relations program for reducing the amount of energy for lighting, heating, and cooling of area residential and commercial buildings.
   h. Evaluate one or more of the following alternative energy sources: (a) wind; (b) solar; (c) geothermal; (d) tidal; (e) underground heat pump; (f) ocean thermal.
   i. Research new energy technology reported in the literature over the past several years.
2. **Energy Consumption in Your Home**. Examine your fuel bills for the past year (or make a yearly projection based on your current monthly bill) to estimate your average monthly and yearly consumption of:
   a. Electricity in kilowatt hours (1 kWh = 3595 kJ)
   b. Natural gas in therms or BTU's (1 therm = 1054 kJ)
   c. Oil in gallons

   On the basis of these estimates for your home, calculate the total average yearly energy consumption in:
   a. Your town or city
   b. Your state
   c. The entire USA
3. **Energy Consumption by Automobiles**. Your car probably runs on gasoline. Determine the weekly consumption of gasoline in gallons by the automobile(s) owned by your family. Now calculate the annual consumption in gallons. Next determine the average yearly consumption of gasoline in:
   a. Your town
   b. Your state
   c. The entire USA

4. A New Source of Crude Oil. A petrologist discovered an underground fossil fuel supply composed of sand, water, and crude oil. Design a series of steps to separate the oil from the mixture. Discuss your proposal with your chemistry teacher. Test your scheme with an artificially prepared mixture in the laboratory. Could you scale up your process so that it could handle large oil spills such as those in the Persian Gulf or off the coast of Alaska?

HISTORY: ON THE HUMAN SIDE

Robert Wilhelm Bunsen (1811-1899), Professor at Kassel, Marburg, and Heidelberg Universities, was known as a great teacher. As a teaching legend at Heidelberg, he taught Adolph Kolbe, Edward Frankland, Victor and Lothar Meyer, Friedrich Beilstein, and Johann Baeyer. Each of these students became a famous chemist in his own right. Bunsen invented the Bunsen burner in 1855 (a similar burner was developed somewhat earlier by Michael Faraday). During research on a highly toxic, unstable compound, he lost an eye during an explosion and twice nearly killed himself through arsenic poisoning. Bunsen and Gustav Kirchhoff invented the spectroscope in 1860 and used it to discover the elements rubidium and cesium.

Lord Kelvin (William Thomson), 1824-1907, matriculated at Glasgow University at the age of 10, and then went on to Cambridge. He assimilated and developed for his book on thermodynamics the pioneering principles of James Joule and Nicolas Sadi Carnot. In his 1852 paper "Dissipation of Mechanical Energy," he set down the law of conservation of energy. Kelvin introduced the concept of absolute zero, recognized the theoretical importance of the absolute temperature scale, and calculated the value of absolute zero from theoretical considerations. With Joule, he demonstrated the Joule-Kelvin effect. A famous quotation of Kelvin's is:

> "I often say that when you can measure what you are speaking about and express it in numbers, you know something about it, but when you cannot express it in numbers, your knowledge of it is of a meager and unsatisfactory kind; it may be the beginning of knowledge, but you have scarcely, in your thought, advanced to the stage of Science, whatever the latter may be."

In 1840, **James Prescott Joule** (1818-1889) discovered a simple law connecting electrical and heat energy. Stated mathematically, $E = I^2Rt$. In 1844, he presented his paper "On the Mechanical Equivalent of Heat" to the Royal Society. Yet both Joule and his father were simple brewers with little formal education. The son carried out his laboratory work in a private laboratory built by his father near their brewery.

Germain **Henri Hess** (1802-1850), although born in Switzerland, was taken to Russia as a child. First a physician, he became a professor of chemistry at the Technological Institute of the University of St. Petersburg. In 1840, by carefully measuring the heats of reaction for many chemical changes, he was able to develop his Law of Constant Heat Summation, known today simply as Hess's Law. The law states that in any chemical reaction, regardless of how many steps there are, the quantity of heat developed in the overall reaction is constant.

Josiah Willard Gibbs (1839-1903) was considered a poor teacher but a brilliant and productive theorist. His early interest in the theory behind James Watt's steam engine governor was responsible for his later interest in thermodynamics. In a series of papers published between 1873 and 1876, he developed virtually all the basic theoretical principles of thermodynamics. His paper "Equilibrium of Heterogeneous Substances" contains the widely used Gibbs "phase rule." In addition, Gibbs developed the vitally important concept of free energy, symbolized by ΔG, that predicts reaction direction.

HUMOR: ON THE FUN SIDE

1. From CHEM 13 NEWS, January 1977, p. 144:

A calorie-counter named Newell
Was confounded for want of a rule,
Till someone implied,
If by 4.184 multiplied,
Each calorie'd turn to a joule.

2. Word Search (see *Appendix* for master copy)

H	Z	B	~~C~~	~~I~~	~~T~~	~~E~~	~~N~~	~~I~~	~~K~~	Y	S	R	T	P	J	H
G	E	Y	U	W	Y	E	H	K	P	H	R	C	C	V	K	N
A	I	J	K	R	B	I	X	O	R	M	L	J	T	W	X	K
V	~~R~~	~~E~~	~~T~~	~~E~~	~~M~~	~~I~~	~~R~~	~~O~~	~~L~~	~~A~~	~~C~~	P	K	G	Y	N
I	O	N	J	S	L	T	X	K	T	W	X	S	U	D	V	H
H	L	C	K	S	N	N	E	Y	U	H	B	E	U	U	N	K
Z	A	M	X	E	E	N	V	A	X	Y	E	C	K	O	D	Q
M	C	V	E	H	H	E	S	O	J	X	K	R	U	W	V	N
R	C	I	C	Q	J	H	X	Y	H	O	O	K	M	K	E	G
E	T	J	L	N	M	M	D	Z	P	W	U	C	G	I	N	O
X	D	Y	Z	F	C	H	F	R	Z	Z	H	L	Z	Y	C	Z
U	P	S	X	~~E~~	~~R~~	~~U~~	~~T~~	~~A~~	~~R~~	~~E~~	~~P~~	~~M~~	~~E~~	~~T~~	L	Y
U	~~T~~	~~H~~	~~E~~	~~R~~	~~M~~	~~O~~	~~C~~	~~H~~	~~E~~	~~M~~	~~I~~	~~S~~	~~T~~	~~R~~	~~Y~~	Y

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. System's measure of disorder or randomness.
2. Device to measure the heat quantity exchanged between a system and its surroundings.
3. SI unit of energy.
4. Measure of the average molecular kinetic energy of a substance.
5. Formulator of a fundamental thermochemical law.
6. Study of the interconversion of heat and other forms of energy in chemical systems.
7. Describing a process that gives off heat to the surroundings.
8. Energy of a system by virtue of its motion.
9. Force acting through a distance.
10. Quantity of heat needed to raise the temperature of one gram of water 1 degree Celsius.

Answers: 1. ENTROPY 2. CALORIMETER 3. JOULE 4. TEMPERATURE 5. HESS 6. THERMOCHEMISTRY 7. EXOTHERMIC 8. KINETIC 9. WORK 10. CALORIE

3. See cartoons at end of the module.

MEDIA

1. *The World of Chemistry* videotape "Number 8: Chemical Bonds" discusses bond strength and energy involvement; "Number 13: The Driving Forces" discusses why reactions occur using heats of reaction. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of

Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.

2. CHEM Study films/videos available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.
    a. *Molecular Motions*. Through animation of molecules, this film illustrates vibrational, rotational and translational kinetic energies. It covers phase changes and simple chemical changes resulting in changes in these three types of energy.
    b. *Vibration of Molecules*. This film describes different vibrational modes in simple molecules. Some of the content may be too difficult for the average college preparatory class.
    c. *Gas Pressure and Molecular Collisions*. This film describes effects of increasing temperature and pressure on gaseous systems using animated models. It is useful in helping students visualize kinetic energy (temperature) and motion of molecules.
3. *"The Laws of Disorder,"* Imperial Chemical Industries, Ltd., ICI Film Library, Thames House North, Millbank, London SW1P 4QG.
4. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    a. *MolVib: Visualizing Molecular Vibrations,* by Daniel Huber. Vol. III C, No. 1, for the Apple Macintosh.
    b. *Microstate*, by Richard York. Vol. V B, No. 2, for the IBM PS/2 PC-compatible computers.
    c. *The Computer-Based Laboratory*, by Daniel Krause. Vol. I A, No. 2, for the Apple II computer.
    d. *Bomb Calorimeter Simulation,* by David J. Olney. Vol. III A, No. 2, for the Apple II computer.
    e. *Notebook: A Column Calculator and Plotter*, by Robert Rittenhouse. Vol. IV B, No. 1, for IBM PS/2 PC-compatible computers.
5. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    a. "An Exothermic Reaction" and "An Endothermic Reaction," two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.). Special Issue 3.
    b. "Thermite Reaction," a chapter on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.). Special Issue 4.
6. *Redox*, by Helen Brooks and David Brooks. Videodisc distributed by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396; (608) 263-2837; (608) 262-0381 (FAX).

EQUIPMENT AND INSTRUMENTATION

The following materials from Project SERAPHIM can be used to build a thermistor probe. Laboratory modules LM 003, LM 004, LM 005, LM 006, LM 015, LM 016, LM 021 (for use with Apple II computers) and LM 022 (for use with IBM PC-compatible computers); LM 023. Published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396; (608) 263-2837; (608) 262-0381 (FAX).

BEST COPY AVAILABLE

Links/Connections

WITHIN CHEMISTRY

1. Molecular motion (translational, rotational, vibrational) and temperature
2. Vibrational kinetic motion and infrared spectroscopy
3. Heats of reaction and potential energy diagrams (used extensively in the *Reaction Rates* module)

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Physics

1. Potential energy, stored energy, heat energy.
2. Gravitational potential energy of a heavy textbook on the floor, work needed to lift it to a table, a position of greater potential energy—all directly analogous to changes in heat energy during a chemical reaction.
3. Heat *vs.* temperature.
4. Heat flow from warmer to cooler surroundings.
5. Calorimetry experiments, especially specific heat determinations of various substances.

Earth Science

1. Snowballs or rocks rolling downhill, losing potential energy as they are converted to kinetic and heat energies.
2. Freezing and/or evaporating of water and the concomitant heat changes.
3. Geothermal energy sources within the earth.

Biology

1. Maintenance of body temperature by combustion reactions within the body.
2. The bombardier beetle provides interesting examples of both exothermic reactions and Hess's Law when it attempts to protect itself from its natural enemies. A pair of glands at the tip of its abdomen has an inner compartment containing hydroquinone, $C_6H_4(OH)_2(aq)$, and hydrogen peroxide, H_2O_2, which are squeezed into an outer compartment containing enzymes (catalysts). A rapid, exothermic reaction occurs, producing quinone, $C_6H_4O_2$, and water, and releasing 204 kJ of heat, according to the equation:

$$C_6H_4(OH)_2(aq) + H_2O_2(aq) \xrightarrow{\text{enzymes}} C_6H_4O_2(aq) + 2H_2O(l) \quad \Delta H = -204\ \text{kJ}$$

The heat raises the solution to its boiling point and the vapor thus produced is sprayed out up to 30 times in quick succession as a hot fine mist with audible sound effects. The Hess's Law example emerges when the above reaction is broken down into its three separate components:

$$\begin{array}{lrcl}
 & 177\ \text{kJ} + C_6H_4(OH)_2(aq) & \rightarrow & C_6H_4O_2(aq) + H_2(g) \\
 & H_2O_2(aq) & \rightarrow & H_2O(l) + {}^1/_2O_2(g) + 95\ \text{kJ} \\
 & H_2(g) + {}^1/_2O_2(g) & \rightarrow & H_2O(l) + 286\ \text{kJ} \\
\hline
\text{Sum:} & C_6H_4(OH)_2(aq) + H_2O_2(aq) & \rightarrow & C_6H_4O_2(aq) + 2H_2O(l) + 204\ \text{kJ}
\end{array}$$

TO THE CONTEMPORARY WORLD

Personal

1. Chemical reactions involving the burning of foods maintain our body temperature and make us fat if we are not careful about what and how much we eat.
2. Attention we all pay to food calories (kilocalories) in our daily diet.
3. Combustion of fuels to warm our homes in winter.
4. Gasoline mileage in our cars related to efficiency of automobile engines.
5. Hot packs and cold packs (used to treat athletic injuries) are the result of chemical reactions also.

Hot pack:

$$CaCl_2(s) \xrightarrow{H_2O} CaCl_2(aq) \quad \Delta H = -82 \text{ kJ/mol } CaCl_2$$

Just 40 g $CaCl_2$ mixed with 100 g water will raise the temperature of the water from 20 °C to 90 °C.

Cold pack:

$$NH_4NO_3(s) \xrightarrow{H_2O} NH_4NO_3(aq) \quad \Delta H = +26.2 \text{ kJ/mol } NH_4NO_3$$

Only 30 g ammonium nitrate mixed with 100 g water will lower the temperature of the water from 20 °C to 0 °C.

Community

1. A field trip to a local power generating plant may be educational, whether a coal-burning, oil-burning, or nuclear plant.
2. Guest speakers from a power plant may be instructive.

SOCIETAL (SCIENCE/TECHNOLOGY/SOCIETY; CURRENT EVENTS)

1. Energy shortages, oil embargo, unfriendly takeover of oil-producing countries, oil wars.
2. Environmental problems caused by our energy needs:
   a. Oil spills
   b. Air pollution, acid rain
   c. Greenhouse effect
3. The energy needed to propel the space shuttle into orbit is equivalent to the energy produced by 23 Hoover dams. The shuttle engines, consisting of three main engines and the two side solid booster rockets, consume 11,000 lb of fuel each second. The total thrust is approximately 44 million horsepower, which is equivalent to 400,000 subcompact cars or 157,000 six-axle diesel locomotives.
4. The solar energy received by the Earth annually is equivalent to 46 times the initial world reserves of coal or 425 times the initial world reserves of crude oil. Only 0.04% of this solar energy is involved in photosynthesis, involving 200 billion tons of CO_2 and releasing 400 billion tons of oxygen.

5. On the last page of his 1923 textbook on thermodynamics, G. N. Lewis expressed concern about the negative ΔG value for the decomposition of NH_4NO_3, especially noting that a mole of solid ammonium nitrate upon heating will decompose into three moles of gas. No one seems to have paid much attention at the time. Ammonium nitrate was a widely used fertilizer, and the storage of large bags filled with it, both in barns and in warehouses, was commonplace. About 25 years later in Texas City, where bags of ammonium nitrate fertilizer were being unloaded on the docks, a fire broke out on one of the ships, and the ship exploded. The fire then spread, and the resulting series of explosions blew up much of Texas City. About 600 people were killed, and almost 4,000 others were injured.

Extensions

1. **Heats of Reaction involving Acids and Bases**
   a. Measure 100 mL of 0.500 M NaOH into a double polystyrene cup. (Nest one cup inside the other.) Place 100 mL of 0.500 M HCl into a beaker. Note and record the initial temperature (°C) of each solution. Rinse and dry both thermometer and graduated cylinder after each use. Take an average of these two initial temperatures as the temperature of the mixture before reaction.

   b. Mix the two solutions with gentle stirring. With the use of a 0.1 °C (or a 1 °C) thermometer, read and record the highest temperature reached as the final temperature, and then calculate the change in temperature, ΔT. Next calculate the amount of heat produced, in joules, assuming the solution has a heat capacity of 4.184 J/g °C. Finally, calculate the quantity of heat produced when you mix one mole of each reactant together.

   c. Repeat the above procedure, using other acid-base combinations: (1) $HNO_3 - NaOH$; (2) $HCl - KOH$; (3) $HNO_3 - KOH$; (4) $H_2SO_4 - NaOH$; (5) $CH_3COOH - NaOH$; (6) $CH_3COOH - NH_3$. What conclusions about the nature of each of these reactions can you make, based on your data?

2. **Heats of Solution of Various Salts**
   a. Prepare two sets of double polystyrene cups, and place 100 mL of water in each. Obtain separately 3.0 g $CaCl_2 \cdot 6\ H_2O$ and 6.0 g NH_4Cl (or NH_4NO_3). Record the initial temperature of the water in each calorimeter to the nearest 0.1 °C (more accurate temperature measurement is desirable). Add the $CaCl_2 \cdot 6\ H_2O$ into the first calorimeter, stir or swirl, and record the highest temperature reached. Repeat with the NH_4Cl in the other calorimeter, recording the lowest temperature reached.

   b. Use the above data to calculate the heat change and heat change per mole, both for the $CaCl_2 \cdot 6\ H_2O$ and the NH_4Cl. Repeat with other salts. Repeat using methanol and 95% ethanol in place of the water.

3. **A Chemist's Map of Pennsylvania**

   Note the energy-related communities on the Pennsylvania map (see *Appendix*): Energy, Boiling Springs, Steam Valley, Crucible, Gastown, Coalport, Coal Hill, and Oil City. Select another state from a detailed map atlas and make a list and/or map of all the sites with names related to energy. Choose one or more of the energy-related sites and try to discover the historical background of the name(s).

4. **Thermal Pollution: A Library Research Project**

   An important environmental concern with regard to nuclear power is thermal pollution. Cooling towers and holding reservoirs for hot water in nuclear plants impact the aqueous environment. Adding hot water to a body of water causes irreversible damage to the organisms of this habitat, killing some flora and fauna and increasing eutrophication. (Idea contributed by Angie Matamoros.)

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The table of covalent bond strengths is taken from pages F177-184.

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41(5), 23-25	Introduction to the Series
41(6), 21-24	The First Law of Thermodynamics
41(7), 36-37	Spontaneity and the Second Law
41(8), 16	Equilibrium: The Limit of Disorder
41(9), 28	Thermal Efficiency and Useful Work
41(10), 24	Factors Affecting Reaction Rates
41(11), 28	Reaction Order and Collision Probability
42(1), 21	How Light Initiates Reactions
42(2), 19	Series Conclusion: Chaos and Evolution

A good essay for students trying to resolve science and religion.

NOTE: This series is said to have included 100 demonstrations.

Appendix

- **Transparency Masters**
  1. Table of Single Covalent Bond Energies (from *Handbook of Chemistry and Physics*)
  2. Enthalpy Changes during Reactions
  3. Schematic Diagram of a Calorimeter
  4. Heats of Reactions between Elements
  5. Enthalpy Changes in the Decomposition and Formation of Water
  6. Dissolving of Solid Ammonium Nitrate (NH_4NO_3): A Typical Endothermic Process
  7. A Chemist's Map of Pennsylvania
  8. Word Search
- **Humor**

Single Covalent Bond Energies

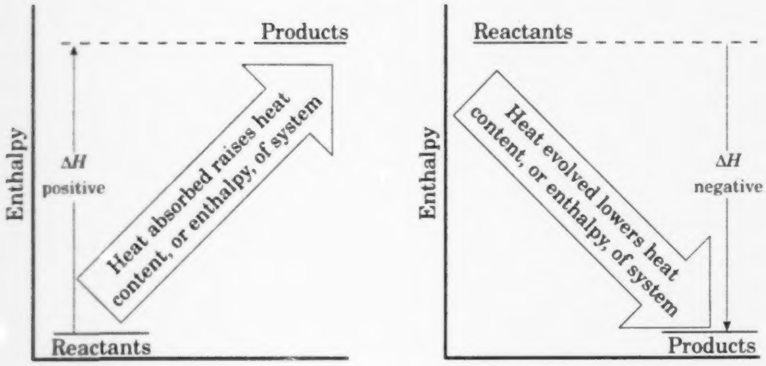
kJ mol^{-1} (kcal mol^{-1})

	F	O	N	Cl	Br	I	C	H	Si
Si	582 (139)	368 (87.9)	335 (80.1)	391 (93)	310 (74)	234 (55)	301 (71.9)	323 (77)	226 (54)
H	566 (135)	467 (111.6)	391 (93)	431 (103)	366 (87.4)	299 (71.5)	416 (99)	436 (104.2)	
C	485 (115.9)	336 (80.3)	285 (68)	327 (78)	285 (68)	213 (50.9)	356 (85)		
I	192.5 (46)	201 (48)		209 (49.9)	180 (43)	151 (36.1)			
Br	238 (56.8)			217 (51.9)	193 (46.1)				
Cl	255 (60.9)	205 (49)	193 (46)	241.7 (57.8)					
N	272 (65)	201 (48)	160 (38)						
O	190 (45)	146 (34.9)							
F	158 (37.8)								

Additional Bond Energies

O=O	498.7 kJ mol^{-1} (118.66 kcal mol^{-1})
N≡N	941.4 kJ mol^{-1} (226.80 kcal mol^{-1})
P_2	489 kJ mol^{-1} (116.70 kcal mol^{-1})
C=C	620 kJ mol^{-1}
C≡C	812 kJ mol^{-1}
C=N	615 kJ mol^{-1}
C≡N	891 kJ mol^{-1}
C=O	781 kJ mol^{-1}
N=N	418 kJ mol^{-1}

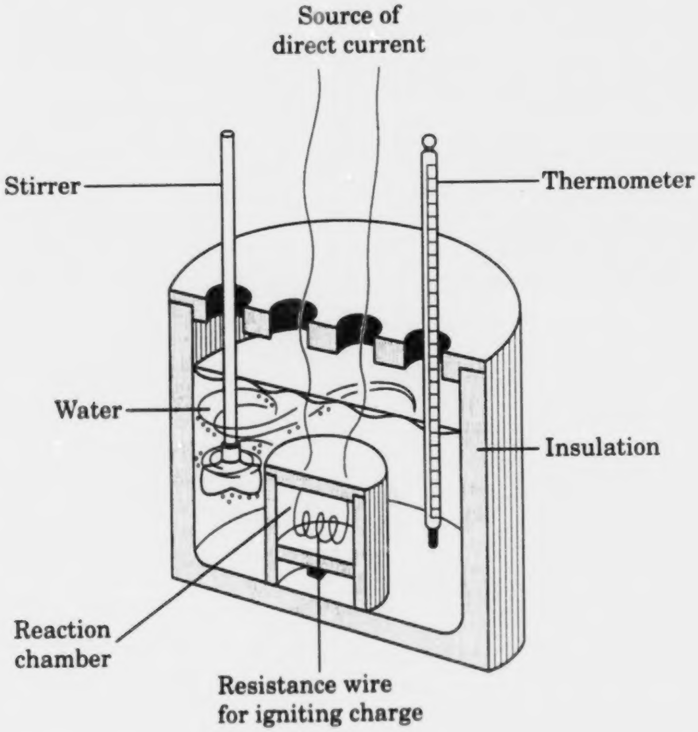
Enthalpy Changes during Reactions


Products
Enthalpy
ΔH
positive
Heat absorbed raises heat content, or enthalpy, of system
Reactants
Reactants
Enthalpy
Heat evolved lowers heat content, or enthalpy, of system
ΔH
negative
Products

Reaction proceeding ⟶
Energy is absorbed
ΔH is positive
(a)

Reaction proceeding ⟶
Energy is evolved
ΔH is negative
(b)

Schematic Diagram of a Calorimeter

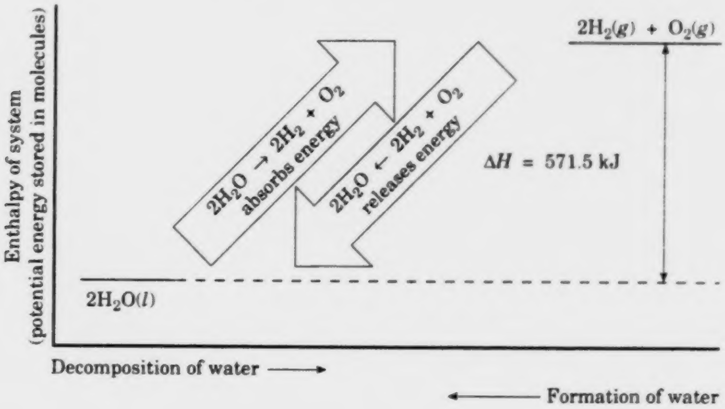

Source of direct current
Stirrer
Thermometer
Water
Insulation
Reaction chamber
Resistance wire for igniting charge

Heats of Reactions between Elements

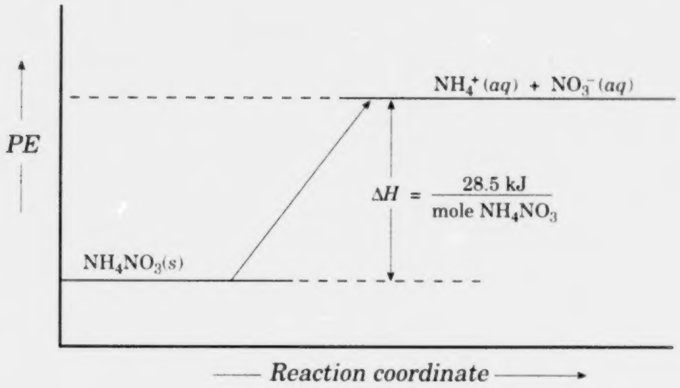
Elements		Compound (product)		Heat of reaction* (kJ/mole of product) ΔH
$H_2(g) + \frac{1}{2}O_2(g)$	$\longrightarrow$	$H_2O(g)$	Water vapor	–241.8
$H_2(g) + \frac{1}{2}O_2(g)$	$\longrightarrow$	$H_2O(l)$	Water	–285.8
$S(s) + O_2(g)$	$\longrightarrow$	$SO_2(g)$	Sulfur dioxide	–297.0
$H_2(g) + S(s) + 2O_2(g)$	$\longrightarrow$	$H_2SO_4(l)$	Sulfuric acid	–811.7
$\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g)$	$\longrightarrow$	$NO(g)$	Nitric oxide	90.4
$\frac{1}{2}N_2(g) + O_2(g)$	$\longrightarrow$	$NO_2(g)$	Nitrogen dioxide	33.9
$\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g)$	$\longrightarrow$	$NH_3(g)$	Ammonia	–46.3
$C(s) + \frac{1}{2}O_2(g)$	$\longrightarrow$	$CO(g)$	Carbon monoxide	–110.5
$C(s) + O_2(g)$	$\longrightarrow$	$CO_2(g)$	Carbon dioxide	–393.3
$C(s) + 2H_2(g)$	$\longrightarrow$	$CH_4(g)$	Methane	–74.5
$2C(s) + 3H_2(g)$	$\longrightarrow$	$C_2H_6(g)$	Ethane	–84.5
$3C(s) + 4H_2(g)$	$\longrightarrow$	$C_3H_8(g)$	Propane	–103.8
$\frac{1}{2}H_2(g) + \frac{1}{2}I_2(g)$	$\longrightarrow$	$HI(g)$	Hydrogen iodide	25.9
$S(s) + \frac{3}{2}O_2(g)$	$\longrightarrow$	$SO_3(g)$	Sulfur trioxide	–394.6

* T = 25 °C, P = 1 atm

Enthalpy Changes in the Decomposition and Formation of Water

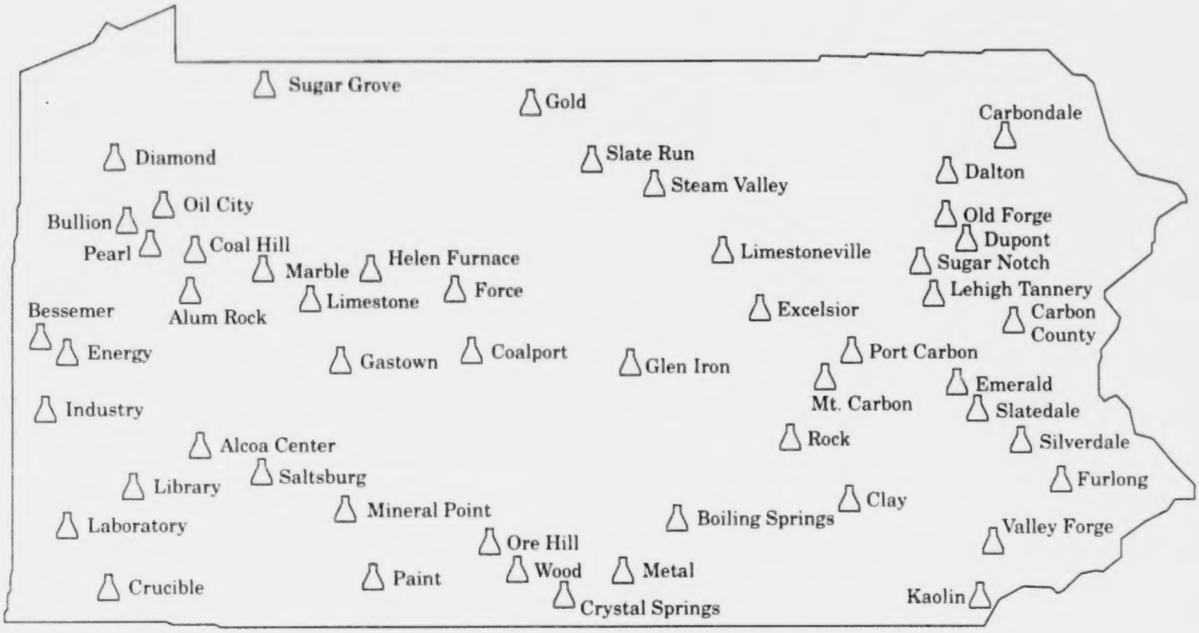

Enthalpy of system
(potential energy stored in molecules)
2H₂(g) + O₂(g)
2H₂O → 2H₂ + O₂
absorbs energy
2H₂O ← 2H₂ + O₂
releases energy
ΔH = 571.5 kJ
2H₂O(l)
Decomposition of water →
← Formation of water

Dissolving of Solid Ammonium Nitrate (NH_4NO_3): A Typical Endothermic Process



$NH_4^+(aq) + NO_3^-(aq)$
PE
$\Delta H = \frac{28.5\ kJ}{mole\ NH_4NO_3}$
$NH_4NO_3(s)$
Reaction coordinate

A Chemist's Map of Pennsylvania


Sugar Grove
Gold
Carbondale
Diamond
Slate Run
Steam Valley
Dalton
Oil City
Bullion
Old Forge
Dupont
Pearl
Coal Hill
Marble
Helen Furnace
Limestoneville
Sugar Notch
Lehigh Tannery
Alum Rock
Limestone
Force
Excelsior
Carbon County
Bessemer
Energy
Gastown
Coalport
Glen Iron
Port Carbon
Mt. Carbon
Emerald
Slatedale
Industry
Rock
Silverdale
Alcoa Center
Library
Saltsburg
Furlong
Clay
Laboratory
Mineral Point
Boiling Springs
Valley Forge
Ore Hill
Wood
Metal
Crucible
Paint
Crystal Springs
Kaolin

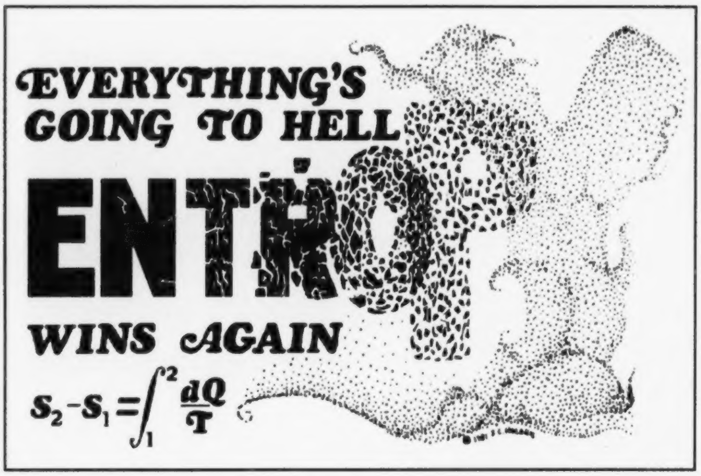
Courtesy of J. J. Bohning, Professor of Chemistry Emeritus, Wilkes University, and Assistant Director, Chemical Heritage Foundation, 3401 Walnut Street, Philadelphia, PA 19104-6228.

Word Search

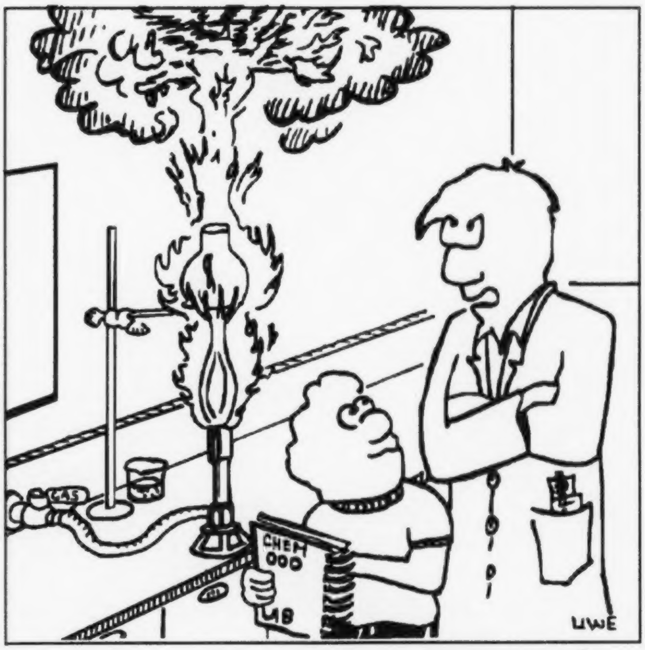
```
H Z B C I T E N I K Y S R T P J H
G E Y U W Y E H K P H R C C V K N
A I J K R B I X O R M L J T W X K
V R E T E M I R O L A C P K G Y N
I O N J S L T X K T W X S U D V H
H L C K S N N E Y U H B E U U N K
Z A M X E E N V A X Y E C K O D Q
M C V E H H E S O J X K R U W V N
R C I C Q J H X Y H O O K M K E G
E T J L N M M D Z P W U C G I N O
X D Y Z F C H F R Z Z H L Z Y C Z
U P S X E R U T A R E P M E T L Y
U T H E R M O C H E M I S T R Y Y
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. System's measure of disorder or randomness.
2. Device to measure the heat quantity exchanged between a system and its surroundings.
3. SI unit of energy.
4. Measure of the average molecular kinetic energy of a substance.
5. Formulator of a fundamental thermochemical law.
6. Study of the interconversion of heat and other forms of energy in chemical systems.
7. Describing a process that gives off heat to the surroundings.
8. Energy of a system by virtue of its motion.
9. Force acting through a distance.
10. Quantity of heat needed to raise the temperature of one gram of water 1 degree Celsius.

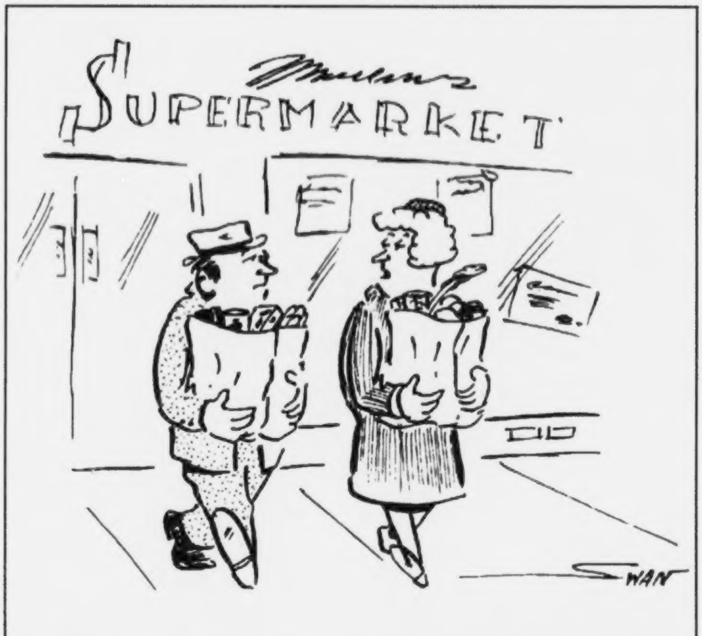

EVERYTHING'S
GOING TO HELL
ENTROPY
WINS AGAIN
$S_2 - S_1 = \int_1^2 \frac{dQ}{T}$

Joyce M Edson, Softwear Unlimited International, P.O. Box 265, Twisp, WA 98856. T-shirts and posters are available from this company. Reprinted with permission.


GAS
CHEM
000
LAB
UWE

YES WILLY, I KNOW IT SAYS TO HEAT BUT!

CHEM 13 NEWS, November 1983, p. 3. Reprinted with permission.



"I'm still not sure what entropy is, but I think it works on money too."

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"NO, NOONKA, IT'S NOT A GOOD THING THAT THE GUYS HAVE DISCOVERED FIRE - BECAUSE NOW THEY'LL EXPECT US TO INVENT COOKING."
Shelly ©91

Cartoon by Shelly Fischman. Reprinted with permission.

"Oh drat! I bet I forgot again to turn off the solar heater!"

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Version 2.0 1994

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ChemSource Project Principal Investigator:
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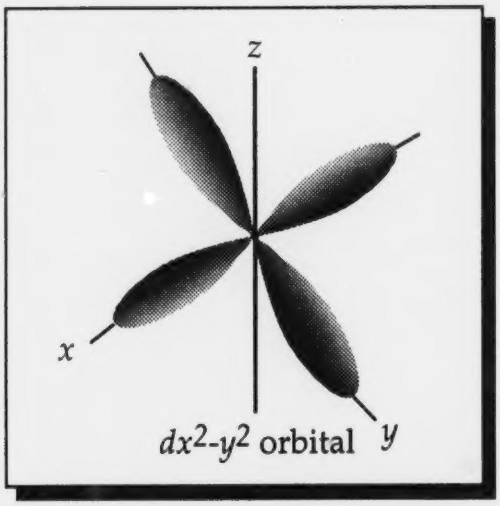
TRANSITION ELEMENTS



ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

The term "transition element" was originally coined to denote the elements in the middle of the Periodic Table that provide a "transition" between the "base formers" on the left and the "acid formers" on the right. Many textbooks refer to all the elements in both the *d*- and the *f*-blocks as transition elements, but some define them as those elements with partially filled *d*- or *f*-orbitals (the d^1-d^9 and f^1-f^{13} elements). SourceBook will refer to the transition elements as those in the *d*- and *f*-blocks.

Most transition elements have several oxidation states, and oxidation-reduction chemistry is important in studying these elements. For example, iron's ability to change readily from 2+ to 3+ is vital to life. This allows Fe^{2+} in cytochrome proteins to oxidize carbohydrates. Both iron oxidation states are ubiquitous in our world. Fe^{3+} is present where oxygen is abundant, while Fe^{2+} is present in oxygen-poor, anaerobic environments.

A pattern can be observed when the number of stable oxidation states is compared to a transition element's position in the Periodic Table. Metals at the left and right sides of the transition element block have the fewest oxidation states (Sc and Zn display only 3+ and 2+, respectively) while those in the middle have the most (Mn has common oxidation states of 2+, 4+, and 7+).

Inorganic compounds of transition elements are often colored. Specific colors are characteristic of particular transition metal ions. These colors help identify ions and sometimes help chemists follow reaction progress. Color adds intrinsic interest to a student's study of transition elements. For example, growing crystals of copper(II) sulfate is both artistically and aesthetically satisfying.

In a representative-element period, such as Period 3, atomic number increase is associated with predictable, distinct variations in element properties. This is not the case for transition elements, which "interrupt" the representative elements within Periods 4, 5, and 6. Periodic variation is still present among transition elements, but it is subtle and requires a more sophisticated interpretation.

This module concentrates on Period 4 transition elements; they are among the most familiar to students and the most accessible for teachers. The suggested laboratory and demonstration activities involve iron, copper, nickel, and manganese. Student activities involving compounds of chromium, nickel, and cobalt—once widely used—are omitted because they are now suspected carcinogens.

An explanation for multiple oxidation states is that electrons in the outermost *s*-orbital and five *d*-orbitals of the transition metal atom are all available for bonding. Because these six orbitals are of nearly identical energies, many electrons are candidates for ionization. For example, Fe can lose two electrons to form Fe^{2+} or it can lose three electrons, forming Fe^{3+}. Electron configurations for such ions are found by removing the outermost *s*-electrons from the neutral atom first, and then removing *d*-electrons as needed. Fe^{2+} is formed when Fe loses its two 4*s*-electrons; Fe^{3+} is formed when Fe loses its two 4*s*-electrons and one 3*d*-electron.

$$\begin{array}{ll} \mathrm{Fe} & [\mathrm{Ar}]4s^2 3d^6 \\ \mathrm{Fe}^{2+} & [\mathrm{Ar}]4s^0 3d^6 \\ \mathrm{Fe}^{3+} & [\mathrm{Ar}]4s^0 3d^5 \end{array}$$

Many transition element orbitals can be involved in covalent bonding. For example, one manganese atom can covalently bond to one, two, or four oxygen atoms producing MnO, MnO_2, or MnO_4^-. In lower oxidation states, the metal ion is often regarded as an individual cation (*e.g.*, Mn^{2+}), although it is often hydrated. In higher states it is covalently bonded within a polyatomic ion (*e.g.*, MnO_4^-). Binary compounds, such as $MnCl_2$ and $CuCl_2$ where each transition element is in a low oxidation state, have bonds with considerable ionic character. As the metal's oxidation number increases the bonds become more covalent in nature.

The presence of color in transition metal compounds is explained by the effect of bonding on *d*-orbitals. Molecules of nonmetals and anions, such as H_2O and Cl^- can act as ligands, bonding to transition metal atoms or ions through an unshared electron pair. The ligand's unshared pair of electrons is shared with the metal ion—thus the bond is regarded as primarily covalent. The entire bonded species, metal plus ligands, is called a coordination complex, or simply a complex. An example is the complex ion $[Cu(H_2O)_4]^{2+}$, which has four water molecules coordinated to the Cu^{2+} ion.

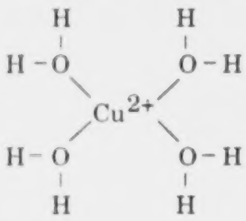
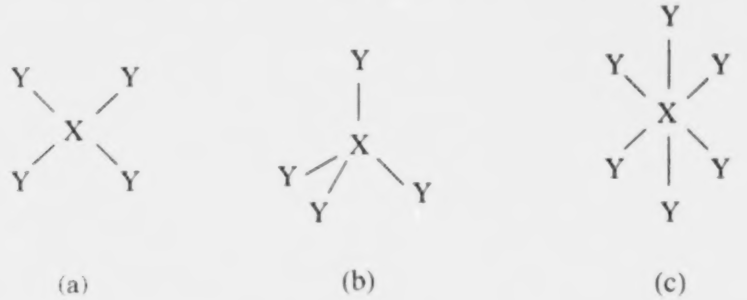


Figure 1. ***$[Cu(H_2O)_4]^{2+}$ complex ion.***

Ligands bond to metals in well-defined numbers and geometries. Over 99% of metal complexes have four or six ligands. Complexes with four ligands have either square planar or tetrahedral geometry, while all complexes with six ligands are octahedral.



(a) (b) (c)

Figure 2. Square planar (a), tetrahedral (b), and octahedral (c) geometries.

$[Cu(H_2O)_4]^{2+}$ is a square planar complex. This complex ion is responsible for the blue color of solid hydrated copper salts, such as $CuSO_4 \cdot 5H_2O$. (The fifth water molecule in this hydrated salt is hydrogen bonded between one coordinated water molecule and the sulfate ion.)

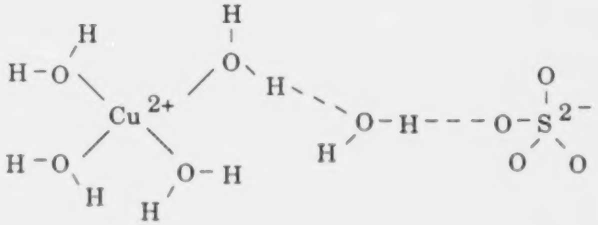


Figure 3. Bonding in $[Cu(H_2O)_5]SO_4$.

The variety of colors exhibited by transition metal compounds is also explained by reference to the metal's *d*-electrons. When all five *d*-orbitals within a subshell participate in bonding with ligands, the *d*-orbitals no longer have the same energy. The size of the energy gap between two sets of *d*-electrons often matches the energy of certain photons of visible light. When an electron moves to a *d*-orbital of higher energy, it absorbs a photon of specific energy. The transmitted light is responsible for the colors. Different metal ions exhibit different colors. Different ligands bonded to the same metal ion generally result in complexes of different colors. Ions with either zero or ten *d*-electrons are colorless. Scandium (Sc^{3+}) has no *d*-electrons to promote; zinc (Zn^{2+}) has no vacant orbital to accommodate any additional electrons. Although not easy, colors of complexes can be predicted using extended forms of this theory.

Similarities in other properties within a transition metal series can be understood by considering structures. Differences in first ionization energies between successive elements within a transition metal series are generally much smaller than those of successive representative elements. An atom's outermost electrons do not feel the full attraction of protons in the nucleus. Some of the "pull" is cancelled by intervening electrons—in effect, the outermost electrons are partially "shielded" from the attractive force of the atom's protons. Even though all electrons are fully attracted to the protons in the nucleus, a given electron is also repelled by other electrons that are momentarily between it and the nucleus. This repulsion is called **shielding**. The net sum of these attractions and repulsions is known as the **effective nuclear charge**.

An example helps to clarify this concept. Selenium, a representative element, has one more proton in its nucleus than does arsenic. It also has an additional electron in a 4*p*-orbital. The added electron shields only 35% of the added proton from the other outer shell electrons. Selenium's outer electrons feel a net increase of 0.65 in effective nuclear charge when compared to the effective nuclear charge in arsenic.

In moving from vanadium to chromium, the added electron enters a 3*d*-orbital, although the highest-energy electrons are in the 4*s*-orbital. The 3*d*-electrons shield 85% of the increased positive charge in the nucleus from the 4*s*-electrons—the 3*d*-electrons reside between the 4*s*-electrons and nucleus about that percent of the time. In general, $n = 3$ electrons are closer to the nucleus than are $n = 4$ electrons. The 4*s*-electrons in chromium atoms experience only 0.15 more effective nuclear charge than do those in vanadium atoms. Effective nuclear charges experienced by outermost electrons increase faster from element to element among representative elements than they do among transition elements. Thus ionization energies for representative elements change more from element to element within a period than do similar values for transition elements. Similar arguments can be used to explain the small changes among atomic radii in the transition series.

The chemistry of transition elements can still be enjoyed without understanding such theoretical nuances. Since trends and relationships are more difficult to predict and explain among transition elements, most students will not benefit from such a theoretical treatment in their first course in chemistry.

PLACE IN THE CURRICULUM

Ideas, demonstrations, and activities from this module can be scattered through the curriculum, since they fit well with other topics (see *Links and Connections* for specific suggestions). If an entire unit on transition elements is taught, material in this module could be used near the end of the year for reinforcement and review of principles. Many activities can also be used to illustrate periodicity and bonding.

CENTRAL CONCEPTS

1. Transition elements are all metals.
2. Many trends in physical properties of transition elements (atomic radius, ionization energy, *etc.*) can be explained by referring to d-orbital electrons.
3. Most transition metals exhibit multiple oxidation states.
4. Many compounds of transition metals are colored.
5. Transition metals and their compounds play an important role in the real world.

RELATED CONCEPTS

1. Electron configurations can be written by assigning electrons to the available orbit of lowest energy.
2. The structure of the Periodic Table is a reflection of the electronic structure of atoms (atomic structure).
3. Bonds are formed when atoms share or transfer electrons (bonding).
4. Oxidation numbers can be assigned to atoms (oxidation-reduction).
5. Redox reactions involve a change in the oxidation number of two or more atoms during reaction (oxidation-reduction).

RELATED SKILLS

1. Recognize trends and patterns in families of the Periodic Table.
2. Calculate gram-gram relationships from a chemical equation.
3. Laboratory manipulations.

PERFORMANCE OBJECTIVES

After completing their study of transition elements, students should be able to:

1. recognize that transition elements are those elements occupying the d- and f-blocks of the Periodic Table.
2. identify color as a property of many compounds of transition elements.
3. identify all transition elements as metals.
4. recognize that most transition elements have multiple oxidation states since d-electrons function as valence electrons.
5. identify transition elements that are important in daily life.
6. recognize that ions of the same element possessing different charges have different properties.

Honors or second year students may also be expected to:

1. interpret some oxidation states in terms of electron structure.
2. identify the ligands and central ion in a complex ion.
3. use laboratory data to illustrate the law of multiple proportions as applied to compounds of transition elements.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Identification of Iron Cations

Introduction

Transition elements form ions with different charges—for example, iron forms Fe^{2+} and Fe^{3+} cations. The properties of these two ions are different.

Purpose

To test for the presence of iron(II) and iron(III) ions with two different reagents and to use observations to design a procedure to test household products.

Safety

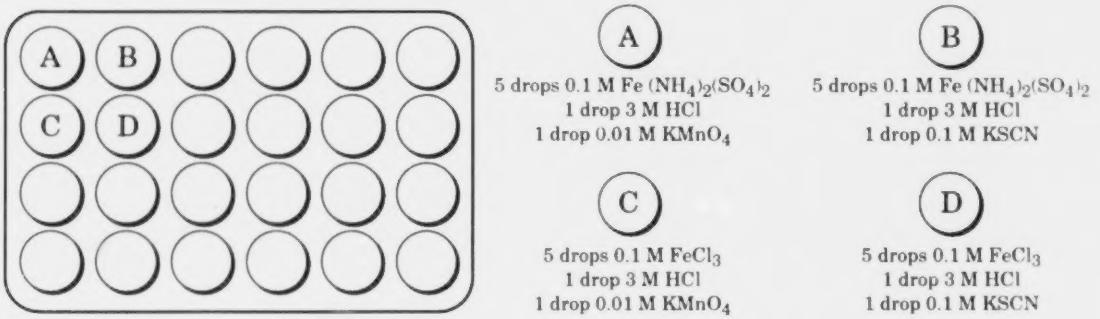
1. Wear protective goggles throughout the laboratory activity.
2. All usual safety precautions should be followed even when working with very small quantities.

Procedure

Prepare a data table and record all of your observations in it.

Part 1

Use a 24-well microtitration plate or six small test-tubes for the activity. Save all the test results until you finish the activity.



A
B
C
D
A
5 drops 0.1 M Fe $(NH_4)_2(SO_4)_2$
1 drop 3 M HCl
1 drop 0.01 M $KMnO_4$
B
5 drops 0.1 M Fe $(NH_4)_2(SO_4)_2$
1 drop 3 M HCl
1 drop 0.1 M KSCN
C
5 drops 0.1 M $FeCl_3$
1 drop 3 M HCl
1 drop 0.01 M $KMnO_4$
D
5 drops 0.1 M $FeCl_3$
1 drop 3 M HCl
1 drop 0.1 M KSCN

1. In each of two wells or test-tubes (labeled A and B) put five drops of 0.1 M $Fe(NH_4)_2(SO_4)_2$ [iron(II) ammonium sulfate] and one drop 3 M HCl (hydrochloric acid). Add one drop 0.01 M $KMnO_4$ (potassium permanganate) to one well and one drop 0.1 M KSCN (potassium thiocyanate) to the other. Record your observations.
2. Repeat Step 1, using 0.1 M $FeCl_3$ [iron(III) chloride] in place of $Fe(NH_4)_2(SO_4)_2$ (Wells C and D).
3. Repeat Step 1, using distilled water in place of $Fe(NH_4)_2(SO_4)_2$.
4. Have your teacher check your data table before starting Part 2.

Part 2

1. Obtain several samples of products you suspect might contain iron. Certain foods, food supplements, plant foods, and rust are examples. If the sample is solid, add 3 M HCl dropwise to make a solution.
2. Write a procedure to test each material for iron(II) and iron(III).
3. Design a data table to record your results.
4. After you obtain your teacher's approval of the proposed procedure, complete your analysis.
5. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

1. Which iron ion was more common among the materials you tested?
2. Can you identify any relationship between the identity of the iron ion and the use of the product?
3. Why was distilled water tested in Part 1?

Implications and Applications

1. How could your procedure be adapted to determine how much iron ion (either Fe^{2+} or Fe^{3+}) is present in a sample?
2. Design a laboratory procedure based on your answer to Question 1.
3. Could your procedure be used to identify both iron(II) and iron(III) in the same sample? Explain.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Identification of Iron Cations

Major Chemical Concept

Compounds of transition elements in different oxidation states exhibit different properties.

Level

General or Basic. (This activity could be adapted for use with honors students by adding more information about electron structures and bonding and/or requiring the *Extension* described later.)

Expected Student Background

Students should be able to recognize that different properties indicate different substances and that properties can be used to identify a substance.

Time

30 to 45 min

Safety

Read *Safety Considerations* in the *Student Version*. Students should be advised that all usual safety precautions should be followed even when working with very small quantities.

Materials (For 24 students working in pairs)

Nonconsumables

12 Microtitration plate or 72 13- x 100-mm test-tubes
Dropper bottles or thin-stem pipets

Consumables

0.1 M Iron(II) ammonium sulfate, $Fe(NH_4)_2(SO_4)_2$ [3.92 g $Fe(NH_4)_2(SO_4)_2 \cdot 6H_2O$ per 100 mL solution]
0.1 M Iron(III) chloride, $FeCl_3$ (2.70 g $FeCl_3 \cdot 6H_2O$ per 100 mL solution; add 4-5 drops 12 M HCl to water before dilution to prevent $Fe(OH)_3$ formation)
0.01 M Potassium permanganate, $KMnO_4$ (0.158 g $KMnO_4$ per 100 mL solution)
0.1 M Potassium thiocyanate, KSCN (0.97 g KSCN per 100 mL solution)
Samples containing iron such as vitamins, raisins, rust, plant food

Advance Preparation

Both 0.1 M $Fe(NH_4)_2(SO_4)_2$ and 0.01 M $KMnO_4$ must be freshly prepared each day—the solutions are not stable. Put solutions in dropper bottles or thin-stem pipettes for each laboratory group.

The easiest way to obtain a variety of household products is to ask students to bring samples from home. This will also add relevancy to the activity. Tap water in some localities contains enough iron cations to give a positive test.

Pre-Laboratory Discussion

This activity could be completed early in the course, even before students have detailed knowledge of electron structure. In that case it would probably be sufficient to emphasize that a change in color generally indicates that a chemical reaction has occurred. The activity might also be used near the end of the course as an opportunity to review what students have learned about chemical analysis. You might ask students to compare electron configurations of an iron atom, and iron(II) and iron(III) ions. With advanced students, discuss the oxidizing properties of $KMnO_4$ and the reducing properties of iron(II) ions.

Teacher-Student Interaction

As you move through the laboratory, insure that students use all their data. A common error is for students to report that Fe(II) did not react with $KMnO_4$ because the solution became nearly colorless. Comparison to the reaction of $KMnO_4$ with water helps. Remind students that disappearance of color as well as appearance of color indicates chemical change. Students will discover that some household products are difficult to test because they possess color. This would be a good time to talk about limitations of tests. If you use microtitration plates be sure that students clean them immediately after use.

Anticipated Student Results

	$KMnO_4$(aq)	KSCN(aq)
Fe(II)	Purple to colorless	Remains colorless
Fe(III)	Slight decolorization over time	Deep red color forms
Distilled water	Remains purple	Remains colorless

The student-designed laboratory procedure should include testing each substance with both reagents. Results can be compared to tests performed on known materials. A sample data sheet is shown below.

	$KMnO_4$(aq)	KSCN(aq)
Vitamin		
Rust		
Plant food		
Raisin		
Cream of Wheat		
Distilled water (control)		

Answers to Data Analysis

1. Answers may vary.
2. Answers may vary. Rust contains iron(III) while vitamins contain iron(II).
3. Distilled water was tested to provide a control. This control allows students to observe the colors of the reagents when no iron(II) or iron(III) is present.

Post-Laboratory Activities

Discuss *Data Analysis* and *Implications and Applications* questions with students. If time permits, students could try out their procedures for the quantitative determination of iron ions. Discuss the concept of the control.

The reaction of Fe^{2+} with $KMnO_4$ can be discussed in terms of the redox reaction, $5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$, where purple MnO_4^- changes to near colorless Mn^{2+}. Reaction of SCN^- with Fe^{3+} produces a 1:1red complex, $Fe(SCN)^{2+}$.

Extensions

1. A Spectronic 20 (or other student spectrophotometer) can be used to determine the amount of a colored substance present in a solution. Activity 18 in

Laboratory Assessment Builds Success gives instructions for such an activity. This activity describes the spectrophotometric analysis of Cu^{2+} ion through conversion to deep blue copper-ammonia complex $[Cu(NH_3)_4]^{2+}$.

2. Potassium hexacyanoferrate(III), $K_3Fe(CN)_6$ reacts with Fe^{2+} to produce $KFe[Fe(CN)_6]$, a deep blue compound sometimes called Prussian blue. Potassium hexacyanoferrate(II), $K_4Fe(CN)_6$ reacts with Fe^{3+} to produce the same compound. Students can devise an additional procedure for testing for the presence of Fe^{2+} and Fe^{3+} using this knowledge.

Assessing Laboratory Learning

1. During the laboratory activity check for the following:
   a. Results of preliminary activity.
   b. Preparation of data table.
   c. Design for testing household products.
2. A substance containing one of the iron cations can be used as an unknown.
3. Paper-pencil questions:
   a. A plant food was tested and the following results were obtained: $KMnO_4$ solution—colorless, KSCN solution—colorless. Which iron ion may have been present? *[Fe^{2+}]*
   b. A dietary supplement was tested with both solutions. The results: $KMnO_4$ remained purple and KSCN changed to red. What is your conclusion about the type of iron in the vitamin? *[Fe^{3+}]*

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: The Amount of Copper in Copper(I) Oxide and Copper(II) Oxide

Introduction

Copper forms two compounds with oxygen. In this activity the two copper-oxygen compounds will be dissolved in hydrochloric acid. The amount of copper in each will be determined by displacing the copper with aluminum wire.

Purpose

To compare the amounts of copper combined with a fixed amount of oxygen in two compounds.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Dispose of all solutions in the container provided by your teacher.

Procedure

1. Label one 100-mL beaker Cu_2O and another CuO. Determine the mass of each beaker.
2. Add about 1 gram Cu_2O to the appropriately labeled beaker and determine the mass of the beaker and contents.
3. Add about 1 gram CuO to the other beaker and determine the mass of this beaker and contents.
4. Add 45 mL 1 M HCl to each beaker and observe. (*CAUTION: Avoid contact of both skin and clothing with hydrochloric acid; it can burn.*)
5. Heat gently with stirring until the oxides dissolve. Copper (II) chloride forms a green solution and the Copper(I) chloride is a white solid. Cool.
6. Make a loop with a handle from each of two pieces of 25-cm aluminum wire. Put one loop of wire in each beaker. Allow the wire to remain in the beaker until there is no further evidence of reaction (*i.e.*, until no more aluminum appears to dissolve).
7. Remove the aluminum wire, with forceps if necessary. Rinse the wire with distilled water from a squeeze bottle. Dislodge any pieces of copper that cling to the wire with a stirring rod. Dispose of the wire as your teacher suggests.
8. Carefully decant the liquid from the solid. Rinse the solid with 10 mL tap water, then with 10 mL distilled water. Pour off as much liquid as possible.
9. Carefully dry the solid by heating the beaker with a burner flame.
10. When the beaker is cool, find the mass of beaker and contents.
11. Heat the beaker and contents for an additional minute. Cool and find the mass. If the mass does not differ significantly from the previous determination, dispose of the copper as your teacher directs.
12. If the mass does differ, repeat Step 11.
13. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

1. Describe the reactants and products.
2. Describe changes in the aluminum wire.
3. Calculate the mass of copper in each compound.
4. Calculate the mass of oxygen in each compound.
5. Calculate the ratio mass Cu/mass O for each compound.
6. Calculate mol Cu/mol O for each compound.
7. Compare the two ratios.
8. Compare the two ratios in terms of atoms.
9. Write chemical equations for all reactions occurring in this activity.

Implications and Applications

1. Explain what effect each of the following would have on the value of the ratio mol Cu/mol O:
   a. The copper was not completely dry before it was weighed.
   b. The reaction was not complete before the aluminum wire was removed from the beaker.
   c. A small piece of aluminum wire remained with the copper.
   d. The copper was not washed thoroughly.
2. Choose symbols to represent copper, aluminum, hydrogen, oxygen, and chlorine atoms. Draw what you would expect to see on an atomic/molecular level at each step of the activity.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: *The Amount of Copper in Copper(I) Oxide and Copper(II) Oxide*

Major Chemical Concept

The Law of Multiple Proportions: If two elements combine to form more than one compound, the amounts of one element combining with a fixed amount of the other element are in the ratio of small whole numbers.

Level

General or Honors

Expected Student Background

Students should be able to:

1. convert from units of grams to moles.
2. use a balance effectively.
3. recognize the need for quantitative separations.
4. perform mole calculations.

Time

45 to 50 min

Safety

Read the *Safety Considerations* in the *Student Version*. Remind students of the hazards involved in using hydrochloric acid.

Materials (For 24 students working in pairs)

Nonconsumables

24 Beakers, 100-mL
12 Burners
12 Spatulas
12 Test-tube holders
12 Forceps
12 Ringstands and rings
12 Wire gauze

Consumables

Copper(I) oxide, Cu_2O, 12 g
Copper(II) oxide, CuO, 12 g
12 Pieces aluminum wire, Al, 18 gauge, 25 cm
1 M Hydrochloric acid, HCl, 540 mL [83 mL conc. (12.0 M) HCl diluted to 1.00 L]
Acetone, 250 mL (*OPTIONAL*: You can hasten the drying process by having students rinse precipitate with ~10 mL acetone. Be sure to use proper safety precautions and disposal procedures. Since acetone is flammable, substitute hot plates for burners if you choose this alternative.)

Advance Preparation

Cut aluminum wire into 25-cm long pieces. It is economical to have students rinse and return pieces of wire remaining after the activity. Be sure the pieces are dry before you store them. Provide spatulas for students to use in transferring the oxides from bottles to beakers.

Pre-Laboratory Discussion

Remind students to make careful measurements and complete separations. It would be a good idea to have students write their mol Cu/mol O ratios on the board. Comparison of several results will help them draw reliable conclusions.

Review the following topics as needed:

1. Data Table Preparation
2. Determining Mass by Difference

(Mass beaker + Contents) − Mass beaker = Mass contents

Mass compound − Mass copper = Mass oxygen

Do not discuss the law of multiple proportions until student data are available. Review the molecular picture of solids and solutions.

Teacher-Student Interaction

During the activity comment on the evidence for change and point out how to determine if the reaction in question is complete. The solutions made by dissolving the copper oxides in acid will be distinctly different. The oxides dissolve somewhat slowly. The "solution" produced by dissolving copper(I) oxide is milky due to the low solubility of copper(I) chloride. The copper(II) chloride solution is blue. Ask students to identify the composition and source of the bubbles that appear when aluminum wire is put in the solution. *[Excess HCl reacts with aluminum to produce hydrogen gas.]* Allow students to make their observations and measurements without focusing on the calculations during the activity. Many students will need help in performing the calculations even if adept at stoichiometry. Comparison of the ratios, if they are put on the board, will help laboratory groups identify sources of error in their results. Students should be able to identify (and reject) ratios that are widely different from those of most other groups.

Anticipated Student Results

Sample Data Table

Mass of Beaker 1	65.26 g
Mass of Beaker 2	64.84 g
Mass of Beaker 1 + Cu_2O	66.26 g
Mass of Beaker 2 + CuO	65.84 g
Mass of Beaker 1 + Cu	66.15 g
Mass of Beaker 2 + Cu	65.64 g
Mass of Cu in Cu_2O	0.89 g
Mass of Cu in CuO	0.80 g

Answers to Data Analysis

1. Copper(I) oxide is a red powder; copper(II) oxide is a black powder. The products in both beakers are a reddish brown solid and a clear liquid.
2. The aluminum wire became thinner.

3. Mass Cu in Beaker 1 (66.15 g - 65.26 g) = 0.89 g

   Mass Cu in Beaker 2 (65.64 g - 64.84 g) = 0.80 g

4. Mass O in Cu_2O (66.26 g - 66.15 g) = 0.11 g

   Mass O in CuO (65.84 g - 65.64 g) = 0.20 g

5. $$\frac{\text{Mass Cu}}{\text{Mass O}} \text{ in } Cu_2O = \frac{0.89\text{ g}}{0.11\text{ g}} = \frac{8}{1}$$

   $$\frac{\text{Mass Cu}}{\text{Mass O}} \text{ in CuO} = \frac{0.80\text{ g}}{0.20\text{ g}} = \frac{4}{1}$$

6. $$\text{Moles Cu in } Cu_2O = 0.89\text{ g Cu} \times \frac{1\text{ mol Cu}}{63.5\text{ g Cu}} = 0.014\text{ mol Cu}$$

   $$\text{Moles O in } Cu_2O = 0.11\text{ g O} \times \frac{1\text{ mol O}}{16.0\text{ g O}} = 0.0069\text{ mol O}$$

   $$\text{Moles Cu in CuO} = 0.80\text{ g Cu} \times \frac{1\text{ mol Cu}}{63.5\text{ g Cu}} = 0.0126\text{ mol Cu}$$

   $$\text{Moles O in CuO} = 0.20\text{ g O} \times \frac{1\text{ mol O}}{16.0\text{ g O}} = 0.012\text{ mol O}$$

   $$\frac{\text{Moles Cu}}{\text{Moles O}} \text{ in } Cu_2O = \frac{0.014\text{ mol Cu}}{0.0069\text{ mol O}} = 2{:}1$$

   $$\frac{\text{Moles Cu}}{\text{Moles O}} \text{ in CuO} = \frac{0.0126\text{ mol Cu}}{0.012\text{ mol O}} = 1{:}1$$

7. The ratio of moles Cu in the two compounds is 2:1

8. The following statements are involved in an atomic-level explanation of this phenomenon:
   a. All atoms of an element are identical.
   b. Atoms combine in whole number ratios.
   c. One mole of any substance has the same number of entities.

9. $$Cu_2O(s) + 2HCl(aq) \rightarrow 2CuCl(s) + H_2O$$

   $$CuO(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O$$

   $$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$

   $$3CuCl(s) + Al(s) \rightarrow AlCl_3(aq) + 3Cu(s)$$

   $$3CuCl_2(aq) + 2Al(s) \rightarrow 2AlCl_3(aq) + 3Cu(s)$$

Implications and Applications

1. a. Ratio too large
   b. Ratio too small
   c. Ratio too large
   d. Ratio too large

2. One possible response is:

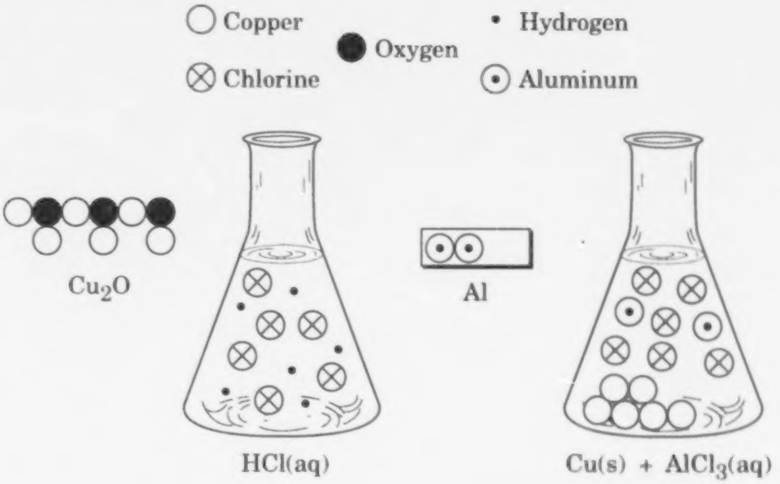


Figure 4. Response for* Implications and Applications *Question 2.

Post-Laboratory Discussion

The post-laboratory discussion can focus on interpretation of student results. The *Implications and Applications* questions provide a starting point. One or two groups will have data that differs widely from the rest of the class. Data on the board are anonymous; you can discuss reasons for discrepancies without humiliating any laboratory group. *Implications and Applications* Question 1 will focus on possible errors. Perhaps students can think of reasons other than those given for lack of precision. Often students with "poor" results know what happened and are willing to offer suggestions. Sometimes results that are far from the norm are due to incorrect calculations. Encourage students to recalculate results that don't seem to fit.

Students may have some difficulty comparing the two ratios. Help them see that we are comparing the amount of copper that combines with a fixed amount of oxygen—one mole. It should be possible from overall class results to conclude that the mole ratio of copper in the two compounds is 2:1.

Once students have mastered the calculations, it is important to focus on matter at the molecular level. Students should share their molecular pictures, since they may have chosen different ways to depict the reactions. Look for consistency in the ratios of copper to oxygen and copper to chlorine.

Assessing Laboratory Learning

1. One chloride of iron is green, the other is orange. What can you predict about the relationship of the mole ratios of iron to chlorine in the two compounds? *[The mole ratios should be small whole numbers because they are two different compounds.]*
2. Would you predict the existence of two different chlorides of magnesium? Explain. *[No. Magnesium is not a transition element and has only one stable oxidation state other than 0.]*
3. Choose symbols to represent manganese and oxygen atoms. Illustrate the formulas for MnO and MnO_2.

• Mn
⊕ O

•⊕ = MnO
⊕•⊕ = MnO_2

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Oxidation States of Manganese

Introduction

Manganese exhibits several oxidation states, and compounds containing manganese have different and interesting colors. Showing them on an overhead projector makes a colorful display. The purpose of this demonstration is to examine oxidation-reduction reactions involving aqueous solutions of manganese ions; to develop a flow chart for the reactions; and to write appropriate equations for the reactions.

Safety

Wear eye protection and an apron throughout the demonstrations. Sulfuric acid is a corrosive substance. Properly dispose of the products.

Materials

Nonconsumables

6 Dropper bottles, 2-4 oz
6 Beakers, 50-mL
Transparency sheet
Overhead projector

Consumables

0.13 M Manganese sulfate, $MnSO_4$ (2.0 g per 100 mL H_2O)
0.0063 M Potassium permanganate, $KMnO_4$ (0.1 g per 100 mL H_2O)
0.16 M Sodium sulfite, Na_2SO_3 (2.0 g per 100 mL H_2O)
6 M Sodium hydroxide, NaOH (24 g per 100 mL H_2O)
50% Sodium hydroxide, NaOH (50 g dissolved in 50 mL water)
6 M Sulfuric acid, H_2SO_4 (33 mL concentrated H_2SO_4 per 100 mL; *CAUTION: Add the acid very slowly to the water.*)

Procedure

1. Place the six small beakers on a transparency sheet on an overhead projector. Using a marking pen, label the beakers *on the underlying transparency*: 2+, 3+, 4+, 5+, 6+, and 7+.
2. Place 3 mL $MnSO_4$ in the 2+ beaker. *(The color will be very pale pink.)*
3. Place 3 mL $MnSO_4$ in the 3+ beaker. Add several droppers of 6 M H_2SO_4. Then add $KMnO_4$ a few drops at a time, swirling the beaker after each addition, until the solution takes on a reddish color.
4. Place 3 mL $MnSO_4$ in the 4+ beaker. Then, without mixing, add 2 or 3 drops of $KMnO_4$. *(Brown spots of MnO_2 will form wherever the drops of potassium permanganate fall.)*
5. Place 3 mL $KMnO_4$ in the 5+ beaker. Add about the same volume of 50% NaOH. Swirl the beaker rapidly until the solution turns blue. *(The color of MnO_4^{3-}.)*
6. Place 3 mL $KMnO_4$ in the 6+ beaker. Add several droppers of 6 M NaOH. Then add several droppers of Na_2SO_3. *(The color changes from purple to green as the MnO_4^- ion is reduced to MnO_4^{2-}, manganate ion.)*
7. Place 3 mL $KMnO_4$ in the 7+ beaker. *(Familiar purple color of MnO_4.)*

[This demonstration could easily be adapted for use as a student activity using 24-well microtitration plates.]

Demonstration 2: Paramagnetism in Transition Metal Compounds

Introduction

Some compounds of transition elements display paramagnetism—the property of being attracted by a magnetic field. This demonstration makes paramagnetism real. It illustrates paramagnetism in a compound of a transition metal and the lack of paramagnetism in a compound such as NaCl.

Materials

Magnet capable of lifting 50 to 100 pounds
Gelatin capsule filled with iron(II) sulfate, $FeSO_4$
Gelatin capsule filled with sodium chloride, NaCl

Procedure

Put the capsules next to the ends of the powerful magnet. As you slowly pull the magnet, the capsule filled with iron(II) sulfate will follow the magnet. The one filled with sodium chloride will not.

Demonstration 3: Nickel One-Pot Reaction

Introduction

In one-pot reactions several reagents are added successively to a single vessel. These are particularly effective when used with transition element compounds—many colors are produced. Precipitations, dissolutions, and reprecipitations are observed in this nickel one-pot reaction. Various reactions of nickel compounds with different reagents are observed. This demonstration offers an opportunity to review descriptive chemistry if you pause between each addition to invite students to predict products and write equations.

Materials

Beaker, 100-mL
Beaker, 1000-mL
Magnetic stirring motor and magnet
Nickel sulfate, $NiSO_4 \cdot 6H_2O$, 3 g
12 M (concentrated) Hydrochloric acid, HCl, 5 mL
2.5 M Sodium carbonate, Na_2CO_3, 23 mL (26.5 g Na_2CO_3 per 100 mL)
6.0 M Nitric acid, HNO_3 12 mL (40 mL 15 M HNO_3 per 100 mL)
0.50 M Sodium hydroxide, NaOH, 110 mL (5 g NaOH per 250 mL)
15 M (concentrated) Ammonia, $NH_3(aq)$, 20 mL
6.0 M Sodium bromide, NaBr, 100 mL (61.8 g NaBr per 100 mL)
3.0 M Sodium hydroxide, NaOH, 100 mL (12 g NaOH per 100 mL)
3.0 M Nitric acid, HNO_3, 250 mL (50 mL 15 M HNO_3 per 250 mL)
1.0% Dimethylglyoxime, 25 mL (1 g dimethylglyoxime per 100 mL water)
25% Ethylenediamine, $H_2NCH_2CH_2NH_2$, 150 mL (37.5 g ethylenediamine per 100 mL water)
1.0 M Sodium sulfide, Na_2S, 15 mL (1.17 g Na_2S per 15 mL)

Safety

Caution is required when handling 12 M HCl, 15 M NH_3, 3 M HNO_3, and 3 M NaOH. All are caustic or corrosive. Sodium sulfide is highly toxic in the present of an acid. It is recommended that this entire one pot reaction be performed in a hood.

Procedure

1. Begin the pot with 3 g of $NiSO_4 \cdot 6H_2O$ in a 50-mL beaker. Add 5 mL 12 M HCl. Pour this solution into 100 mL distilled water in a 1000-mL beaker on a magnetic stirrer. Rinse any nickel sulfate crystals that remain into the larger beaker.
2. Add 23 mL 2.5 M sodium carbonate.
3. Dissolve the precipitate with 12 mL 6.0 M nitric acid.
4. Add 110 mL 0.50 M sodium hydroxide.
5. Add 20 mL 15 M ammonia.
6. Add 100 mL 6.0 M sodium bromide. Allow enough time for the precipitate to form.
7. Add 100 mL 3.0 M sodium hydroxide.
8. Acidify the pot with 250 mL 3.0 M nitric acid.
9. Add 15 mL 25% ethylenediamine.
10. Add 25 mL 1.0% dimethylglyoxime solution.
11. Add 15 mL 1.0 M sodium sulfide.

Teacher Notes

Colors of Nickel Complexes

Formula	**Color**	**Geometry**
$NiCl_4^{2-}$	Yellow	Tetrahedral
$Ni(H_2O)_6^{2+}$	Green	Octahedral
$Ni(NH_3)_6^{2+}$	Blue	Octahedral
*$Ni(en)_3^{2+}$	Purple	Octahedral
*$Ni(dmg)_2$	Red	Square planar

*en = ethylenediamine, dmg = dimethylglyoxime

Equations for reactions:

1a. $Ni^{2+}(aq) + 4Cl^-(aq) \rightarrow NiCl_4^{2-}(aq)$

1b. $NiCl_4^-(aq) \rightarrow Ni^{2+}(aq) + 4Cl^-(aq)$

2. $Ni^{2+}(aq) + CO_3^{2-}(aq) \rightarrow NiCO_3(s)$

3. $NiCO_3(s) + 2H^+(aq) \rightarrow Ni^{2+}(aq) + H_2O(l) + CO_2(g)$

4. $Ni^{2+}(aq) + 2OH^-(aq) \rightarrow Ni(OH)_2(s)$

5. $Ni(OH)_2(s) + 6NH_3(aq) \rightarrow Ni(NH_3)_6^{2+}(aq) + 2OH^-(aq)$

6. $Ni(NH_3)_6^{2+}(aq) + 2Br^-(aq) \rightarrow Ni(NH_3)_6Br_2(s)$

7. $Ni(NH_3)_6Br_2(s) + 2OH^-(aq) \rightarrow Ni(OH)_2(s) + 2Br^-(aq) + 6NH_3(aq)$

8. $Ni(OH)_2(s) + 2H^+(aq) \rightarrow Ni^{2+}(aq) + 2H_2O(l)$

9. $Ni^{2+}(aq) + 3en(aq) \rightarrow Ni(en)_3^{2+}(aq)$

10. $Ni(en)_3^{2+}(aq) + 2dmg^-(aq) \rightarrow Ni(dmg)_2(s) + 3en(aq)$

11. $Ni(dmg)_2(s) + S^{2-}(aq) \rightarrow NiS(s) + 2dmg^-(aq)$

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What type of electron configuration do the transition elements have in common? *[All have one or more electrons in their* d*-sublevel. Only the zinc group has a completely filled* d*-sublevel.]*
2. Transition elements exhibit a high frequency of multiple oxidation states. How can this be explained? *[The d-electrons as well as the s-electrons can be used as bonding electrons.]*
3. Compounds of transition elements often exhibit color. To what can this color be attibuted? *[If the transition element forms a compound in which its five d-orbitals do not all have the same energy, the result is an energy gap that matches the energy of visible light photons. When white light strikes the compound, those photons are absorbed. The color we see is the result of the colors that are not absorbed.]*
4. How is a gemstone different from a mineral? *[A gemstone forms when a particular transition element atom is incorporated as an impurity in a parent mineral. For example, ruby results from* Cr^{3+} *ions replacing some of the* Al^{3+} *ions in the mineral corundum,* Al_2O_3*. A mineral is a naturally occurring chemical compound.]*
5. Consider the periodic relationships in the transition series following:

	Element	1st Ionization Energy, kJ/mol	Atomic Radius, pm
3rd period	Na	496	157
	Mg	738	136
	Al	577	125
	Si	787	117
	P	1012	110
	S	1000	104
	Cl	1251	99
4th period	K	419	203
	Ca	590	174
	Sc	631	144
	Ti	658	132
	V	650	122
	Cr	653	117
	Mn	717	117
	Fe	759	117
	Co	760	116
	Ni	737	115
	Cu	745	117
	Zn	906	125
	Ga	579	125
	Ge	762	122
	As	947	121
	Se	941	117
	Br	1140	114
	Kr		

Using data in the table, prepare the following graphs:

a. Atomic Radii *vs.* Atomic Number

b. Ionization Energy *vs.* Atomic Number

For elements in Period 4, plot number of *d*-electrons *vs.* atomic number.

Answer these questions about the graphs:

1. How do the graphs for atomic radii differ for the two periods? *[Period 4 has a group of atoms with similar radii while Period 3 shows a general decrease.]*
2. How do the graphs for ionization energy differ for the two periods? *[Period 4 increases and then remains fairly constant while Period 3 shows a general increase.]*
3. How do these trends in ionization energy and atomic radius help in predicting properties of these elements? *[Those properties dependent upon effective nuclear charge will remain fairly constant throughout the* d*-block. Such properties, in addition to ionization energy and atomic radius, are ionic radius and density. By contrast, some properties such as paramagnetism, melting point and boiling point, depend upon the number of unpaired electrons and vary widely across the d-block. You may wish to use this fact for extension studies.]*
4. Use the graph of number of *d*-electrons *vs.* atomic number to explain the following oxidation states of Period 4 elements. (These represent only a few oxidation states for these elements. Some can only be predicted with considerably more information.)

a. Ca 2^+ only   b. Sc 3^+ c. Mn 7^+

d. Cr 6^+ e. Fe^{2+} and Fe^{3+} f. Zn 2^+ only

[All period 4 elements initially lose their 4s-electrons upon oxidation. Period 4 elements with a partially filled d*-subshell can also lose* d*-electrons up to the number contained in the subshell.]*

a. $Ca \rightarrow Ca^{2+} + 2e^- \ [Ar]\,3d^0\,4s^2 \rightarrow [Ar]\,3d^0\,4s^0$

b. $Sc \rightarrow Sc^{3+} + 3e^- \ [Ar]\,3d^1\,4s^2 \rightarrow [Ar]\,3d^0\,4s^0$

c. $Mn \rightarrow Mn^{7+} + 7e^- \ [Ar]\,3d^5\,4s^2 \rightarrow [Ar]\,3d^0\,4s^0$

d. $Cr \rightarrow Cr^{6+} + 6e^- \ [Ar]\,3d^4\,4s^2 \rightarrow [Ar]\,3d^0\,4s^0$

e. $Fe \rightarrow Fe^{2+} + 2e^- \ [Ar]\,3d^6\,4s^2 \rightarrow [Ar]\,3d^6\,4s^0$

$Fe \rightarrow Fe^{3+} + 3e^- \ [Ar]\,3d^6\,4s^2 \rightarrow [Ar]\,3d^5\,4s^0$

f. $Zn \rightarrow Zn^{2+} + 2e^- \ [Ar]\,3d^{10}\,4s^2 \rightarrow [Ar]\,3d^{10}\,4s^0$

Counterintuitive Examples

1. Compounds with formulas such as Fe_3O_4 and $MnO_{1.85}$ are examples of nonstoichiometric compounds. They are generally compounds where transition elements exist in two or more oxidation states.
2. Manganese (IV) oxide is an example of a compound that can act as either an oxidizing agent or a reducing agent. The relevant half reactions are:

Oxidizing agent: $MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$

Reducing agent: $MnO_2(s) + 2H_2O(l) \rightarrow MnO_4^-(aq) + 4H^+(aq) + 3e^-$

TIPS FOR THE TEACHER

Language of Chemistry

coordination complex species made up of metallic ion and the ligands that surround it.

covalent bonding bonding pattern in which a bond consists of one or more pairs of shared electrons.

diamagnetism weak repulsion by a magnetic field.

effective nuclear charge net sum of all the attractive and repulsive forces experienced by an electron in an atom.

electron configuration shorthand notation describing the distribution of electrons among the energy levels of an atom.

ferromagnetism ability of a substance to become permanently magnetized by exposure to an external magnetic field.

ionic bonding bonding pattern characterized by the transfer of electrons from one bonding atom to the other which results in the formation of ions.

ionization energy energy required to remove an electron from an atom.

ligands molecules or ions that surround a metal ion in a complex.

octahedral configuration shape in which six species are equally spaced around a central species.

oxidation loss of electrons by a species.

oxidation state the apparent charge on an atom if the electrons are assigned according to established rules.

paramagnetism attraction toward a magnetic field (stronger than diamagnetism, weaker than ferromagnetism).

reduction gain of electrons by a species.

refractory high melting.

representative elements elements of the *s*- and *p*-blocks.

square planar configuration arrangement in which four species are at the corners of a square surrounding a fifth species in the center.

transition elements elements of the *d*- and *f*-blocks.

valence electrons outer energy level electrons involved in bonding, which include the *d*-electrons for the *d*-block transition elements.

Pattern Recognition

A Comparison of Representative and Transition Elements

Representative Elements

About half are metals.
Atomic size increases from F to Cs.
Vary in electronegativity from 0.7 (Cs) to 4.0 (F).
Simple ions in water are usually colorless.
Metals usually exhibit only one oxidation state.
Valence electrons are all in the outermost shell.
Oxides are normal (metal oxides are basic, nonmetal oxides are acidic).
Simple hydrides are normal in composition and properties.
Not usually associated with catalytic activity.
Not usually magnetic (maximum number of unpaired electrons is only three).
Metallic activity increases from top to bottom within a group.
Electronic structures of succeeding elements differ by one electron in the outer shell.

Transition Elements

All are metals.
Atomic sizes about the same within a period.
Electronegativity relatively constant (about 1.7).
Simple ions in water usually colored.
Metals can usually have a number of different oxidation states.
Valence electrons are in more than one shell.
Oxides are abnormal, some basic, some acidic, some refractory (ceramic compounds with low thermal conductivity able to withstand very high temperatures).
Hydrides are usually abnormal in composition and properties.
Most exhibit pronounced catalytic activity.
Often paramagnetic, sometimes ferromagnetic (as many as seven unpaired electrons in *f*-block).
Metallic activity decreases from top to bottom within a group.
Generally electronic structures of succeeding elements differ by one electron in an inner shell while the number of outer shell electrons is constant.

Common Student Misconceptions

1. **"All metals are magnetic."**

   Demonstration 2 can correct this misconception.

2. **"All compounds of transition elements are colored."**

   Examples of zinc and/or scandium compounds can be used to disprove this idea. For honors students, the explanation for lack of color is that zinc has a filled *d*-shell and scandium has no *d*-electrons.

HISTORY: ON THE HUMAN SIDE

In 1893, Alfred Werner, a 26-year-old Swiss chemist described the nature of complex ions by proposing that ammonia, chloride ion and similar ligands could be strongly attracted to metal ions to form complex ions. His predictions were based on the reactions of cobalt complexes with silver nitrate. From observing quantities of silver nitrate required to precipitate chloride ions, he suggested that chloride ions could be bonded ionically or by what he called "secondary bonds." He even predicted the geometries of the compounds that were confirmed many years later by x-ray crystallography. His accomplishments are even more astounding when it is realized that his work preceded Lewis bonding theory by 23 years. He received the Nobel Prize in chemistry in 1913.

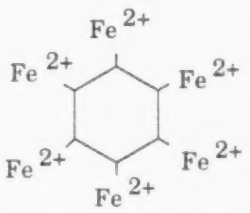
HUMOR: ON THE FUN SIDE

1. Students enjoy, even though they will groan loudly, these puns based on transition elements.

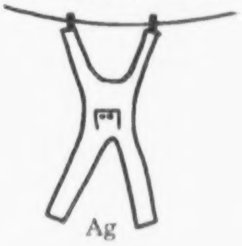
HIOAg **Hi O Silver!**

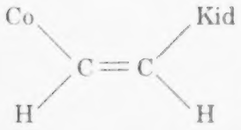
Ferrous Wheel

[Explain the now obsolete nomenclature of Fe^{2+} (ferrous) and Fe^{3+} (ferric) ions.]



LONG JOHN SILVER





Cisco Kid

[Explain about *cis*- and *trans*-geometric isomers. See *Organic Chemistry* module.]

2. **New scientific fiction book**: "Where's Mercury Found?" by H. G. Wells.

   CHEM 13 NEWS, September 1979, p. 11

3. **Nomenclature riddles**:

   a. What is CFe_2O_3? **Answer**: Crust

   CHEM 13 NEWS, November 1976, p. 1107

   b. Who is Ni(II) + Cu + Ni(II) [short form]? **Answer**: Nicolaus Copernicus

   c. What is Fe^{2+} $C_{10}H_{22}$? **Answer**: Ferris Wheel

   CHEM 13 NEWS, September 1974, p. 783

4. **LANTHANIDE MNEMONIC**

   Lettuce can provide nourishing protein, so everyone
   La Ce Pr Nd Pr Sm Eu

   gets tender, delicious, wholesome, enriched, young lettuce.
   Gd Tb Dy Ho Er Yb Lu.

   CHEM 13 NEWS, May 1981, p. 9

5. **CHEMTOON**

   Fe Cr Let's get a shine on!

   CHEM 13 NEWS, October 1983, p. 16

5. **Roaming Through the Elements—How Many Can You Name?**

   a. Eight elements mentioned in the Bible.
   b. Three elements named after women.
   c. Eight elements named after countries.
   d. Seven elements named after famous scientists.
   e. Ten elements named after celestial bodies.
   f. Five elements whose names are of German origin.
   g. One element whose name is of Spanish origin; one Swedish.
   h. Three elements named after continents or states.
   i. Eleven elements named after districts, cities, towns, or villages.
   j. Twenty-seven elements whose names or symbols are of Latin origin.
   k. Thirty elements whose names are of Greek origin.

   The answers are:

   a. Au, Ag, Fe, Sn, Pb, Cu, C (as diamonds), S (as brimstone)

   b. Cm (Marie Curie); Nb (Niobe, daughter of Tantalus); V (Vanadis, Scandinavian goddess)

   c. Fr, Ga (Gallia, Latin name for France), Ge, Po, Ru (Ruthenia, Latin name for Russia), Sc (Scandia, Latin name for Scandinavia), Tm (Thulium, ancient name for Scandinavia), In.

   d. Cm, Es, Fm, Gd, Lr, No, Md

   e. Ce (Ceres, an asteroid); He (Helios, Greek name for the Sun); Hg; Np; Pd (Pallas, an asteroid); P (ancient name for Venus when appearing before sunrise); Pu; Se (Selene, Greek name for the Moon); Te (Tellus, Latin name for the earth); U

   f. Bi (Latinized form of Wissmuth, Weisse Masse = white mass); Co (Kobold = goblin or evil spirit); Ni (from Kupfernickel = Old Nick's copper); Zn; Sn

g. Pt (Platina = silver); W (Tung = heavy, sten = stone)

h. Am, Eu, Cf

i. Bk (home of the University of California); Er, Tb, Yb, Y (all for Ytterby, a village in Sweden); Hf (Hafnium, the Latin name for Copenhagen); Ho (Holmium, the Latin name for Stockholm); Lu (Lutetia, an old name for Paris); Mg (Magnesia, a district in Thessaly); Re (Rhenium, the Latin name for Rhine); Sr (Strontian, a town in Scotland)

j. Sb (stibium = mark), Ca (Calx = lime), Cs (caesium = sky blue), Cu (cuprum = from Cyprus), F (fluere = flux or flow), Ga (Gallia = France), Au (aurum = shining dawn), Hf (Hafnia = Copenhagen), Ho (Holmia = Stockholm), Ir (iris = rainbow), Fe (ferrum, Latin name for iron), Pb (plumbum, Latin for lead), Mn (magnes = magnet), Hg (hydrargyrum = liquid silver), K (kalium = alkali), Ra (radius = ray), Rn (radon from radium), Re (Rhenus = Rhine), Rb (rubidius = deepest red), Ru (Ruthenia = Russia), Sc (Scandia = Scandinavia), Si (silex = flint), Ag (argentum, Latin for silver), Na (natrium, Latin for sodium), Te (tellus = earth), Sn (stannum, Latin for tin), Ti (titans = first men on earth

k. Ac (actinos = beam or ray), Ar (argon = inactive), As (arsenicon = yellow orpiment), At (astatos = unstable), Ba (barys = heavy), Br (bromos = stench), Cl (chloros = greenish yellow), Cr (chroma = color), Dy (dysprositos = hard to get at), He (helios = sun), H (hydro = water, genes = forming), I (iodes = violet), Kr (kryptos = hidden), La (lanthanein = to lie hidden), Li (lithos = stone), Mo (molybdos = lead), Nd (neo = new, didymos = twin), Ne (neos = new), N (nitron = native soda, genes = forming), Os (osme = smell), O (oxys = sharp, acid, genes = forming), P (phosphoros = light-bearing), Pr (praseos = green, didymos = twin), Pa (protos = first), Rh (rhodos = rose), Se (selene = moon), Ta (Tantalus = father of Niobe), Tc (technekos = artificial), Tl (thallos = green shoot or twig), Xe (xenos = stranger)

CHEM 13 NEWS, October 1980, p. 5

6. **CHEMISTRY CAN BE FUN**

Fill in each blank with the name of the chemical element from the alphabetized list. One element is used twice.

The _1_ and jewels _2_. That _3_ must have stolen them. Call the _4_ before he has had a chance to _5_ or _6_. There should be a _7_ his place before he has got _8_. When they break _9_ the _10_ will _11_ and _12_ to give them back. If he doesn't, I wouldn't give a plugged _13_ for his chances. He will be behind _14_ bars in no time. Maybe that will _15_.

Alphabetized list:

1. argon 2. barium 3. cesium 4. copper 5. curium 6. dysprosium 7. gold 8. indium 9. iridium 10. iron 11. nickel 12. radon 13. silicon 14. tellurium

CHEM 13 NEWS, April 1977, p. 14

7. **Transition Metals Trivial Pursuit** (Adapted from CHEM 13 NEWS, September, 1983, p. 11)

Choose the symbol of one of the following elements that, either in the free or in the combined state, best satisfies each clue: Ti, Cr, Fe, Ni, Zr, Ce, Pr, Nd, Sm, Tb, W, Re, Os, Pt

a. The hardest metal.

b. Lanthanide discovered by Mosander in 1842.

c. Strongest reductant of the rare earth dipositive ions.

d. Green twin.

e. Relative of manganese

f. Ferberite is one of its ores.

g. An oxide of this metal is yellow when cold, orange when hot.

h. A dispositive rare earth insoluble sulfate.

i. The sulfuric acid solution of the peroxidized form of this tetrapositive metal is bleached by Hf.

j. An aqueous mixture of the salts of neodymium and this rare earth is colorless, analogous to Ni - Co mixture.

k. Didymium glass used in glassblowers' goggles contains a mixture of the oxides of praseodymium and this rare earth.

l. An oxide of this metal is white when cold, yellow when hot.

m. The compounds of this element exhibit disproportionation to a greater extent than any other element.

n. This element is quantitatively precipitated by phosphoric acid in (1 : 9) sulfuric acid solution containing H_2O_2.

o. An oxide of this metal, sometimes erroneously named as an acid, is used as a histological reagent.

p. The highest melting metal.

q. In the dipositive state it is quantitatively precipitated in slightly acid solution by dimethylglyoxime as fine blue crystals with a metallic bronze luster.

Answers:

a. Cr	d. Pr	g. W	j. Pr	m. Re	p. W
b. Tb	e. Re	h. Sm	k. Nd	n. Zr	q. Pt
c. Sm	f. W	i. Ti	l. Ti	o. Os	

8. Word Search (see *Appendix* for master copy)

~~D~~	~~N~~	~~A~~	~~G~~	~~I~~	~~L~~	O	X	M	X	O	B	X	L	K	N	S
B	F	O	D	H	Y	C	H	N	~~P~~	~~U~~	~~R~~	~~P~~	~~L~~	~~E~~	O	J
I	W	U	L	U	G	U	H	D	T	L	V	D	M	S	I	P
Y	N	M	I	W	O	N	D	R	A	O	Z	T	C	T	T	L
N	R	Z	O	P	N	R	O	R	O	D	O	D	Q	U	A	H
L	E	P	N	Y	G	R	D	I	N	M	W	N	H	M	N	U
E	P	N	I	U	D	E	W	M	T	L	I	F	R	G	I	Z
W	P	S	Z	L	H	N	U	W	T	I	E	U	X	Z	D	K
K	O	I	A	A	U	R	X	D	Y	M	S	W	M	I	R	I
I	C	V	T	K	E	E	A	T	J	C	L	N	D	X	O	G
G	N	C	I	N	D	W	R	R	S	D	E	S	A	H	O	Y
F	O	N	O	V	M	K	C	V	M	K	S	U	C	R	C	K
M	V	R	N	R	A	J	~~N~~	~~O~~	~~R~~	~~I~~	Z	C	G	H	T	C

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Molecule or ion surrounding a metal ion in a complex.
2. Adjective for the *d*- and *f*-block elements.
3. Shape associated with six species equally spaced around a central species.
4. Transition element identified by electron configuration [Ar]$3d^6 4s^2$.
5. Color of aqueous solution of permanganate ion.
6. In addition to tin, bronze statues contain this transition metal that turns them green upon aging.
7. Transition element responsible for the red color in rubies.
8. Type of energy required to remove an electron from an atom.
9. Type of complex composed of a metallic ion and its surrounding ligands.
10. Swiss chemist who predicted geometries of complex ions.

Answers: 1. LIGAND 2. TRANSITION 3. OCTAHEDRAL 4. IRON 5. PURPLE 6. COPPER 7. CHROMIUM 8. IONIZATION 9. COORDINATION 10. WERNER

9. See cartoons at end of module.

MEDIA

1. The *World of Chemistry* videotape "Number 19: Metals." World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224. A secondary school version of this series is available from WINGS for Learning/SUNBURST, 101 Castleton Street, Pleasantville, NY 10570; (914) 747-3310; (800) 321-7511; (914) 747-4109 (FAX).
2. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   a. *KC? Discoverer with Knowledgeable Counselor,* by Daniel Cabrol, John W. Moore and Robert C. Rittenhouse. Special Issue 2, for IBM PS/2, PC compatible computers.
   b. *KC? Discoverer: Exploring the Properties of the Chemical Elements,* by Aw Feng and John W. Moore. Vol. I B, No. 1, for IBM PS/2, PC compatible computers.
   c. *KC? Discoverer?,* by Michael Liebl, Vol. IV A, No. 2, for all Apple II computers.
   d. *The Periodic Table Stack,* by Michael Farris. Vol. I C, No. 1, for the Apple Macintosh.
   e. *The Periodic Table* (Toolbook), by Paul F. Schatz, John C. Kotz and John W. Moore, in press. For Windows running on IBM PS/2 and PC-compatible computers.

   f. *Notebook: A Column Calculator and Plotter*, by Robert Rittenhouse. Vol. IV B, No. 1, for IBM PS/2 PC-compatible computers.
   g. *The Computer-Based Laboratory*, by Daniel Krause. Vol. I A, No. 2, for the Apple II computer.

3. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
   a. For the Apple II computer running on ProDOS: AR 601, AR 1201.
   b. For the Apple II computer: AP 301, AP 608, AP 1201; Laboratory modules LM 002, LM 003, LM 004, LM 005, LM 006, LM 012, LM 015, LM 016, LM 021 (for the Project SERAPHIM Blocktronic and Thermistor).
   c. For IBM PCs and PC-compatibles: PC 3301, PC 3401, PC 4201; Laboratory modules LM 002, LM 003, LM 004, LM 005, LM 006, LM 012, LM 015, LM 016 (for the Project SERAPHIM Blocktronic and Thermistor).
4. Videodisc published by *JCE: Software,* a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

The Periodic Table Videodisc (single side, 30 min.). Special Issue 1.

5. *The Periodic Table*, Minnesota Educational Computing Consortium, St. Paul, MN (MECC; a computer program, database, and related instructional materials).
6. *Doing Chemistry* (some reactions on this videodisc series are reactions of transition elements). Available from American Chemical Society, 1155 Sixteenth Street, N.W., Washington, DC 20036.

EQUIPMENT AND INSTRUMENTATION

1. Computer and software to plot graphs and use the MECC program (see *Media*).
2. Colorimeter, spectrophotometer, or computer interface (Blocktronic).

Links/Connections

WITHIN CHEMISTRY

1. **Reactions Module**. Many reactions in the laboratory activities and demonstrations involve transition metal compounds.
2. **Redox Module.** One reaction in "Identification of Iron Ions" laboratory activity is an example of a redox reaction:

$$5Fe^{2+}(aq) + MnO_4^{-}(aq) + 8H^{+}(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

3. **Periodicity Module.** The ionization energy and radii plotting exercise for all Period 4 elements can be used as an extension during the discussion of periodicity. The graphs will make the point—quite strongly—that changes are less pronounced among transition metals.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Biology

Hemoglobin binds and transports O_2 to cells with the help of a transition metal complex. The O_2 actually is the sixth ligand in a Fe^{2+} complex. The other five ligands include N atoms. At the cellular level, O_2 transfers to myoglobin, which stores it until it is needed by the muscle. Myoglobin also binds oxygen as the sixth ligand of a Fe^{2+} complex. Poisoning by carbon monoxide occurs because CO competes with O_2 for bonding positions to the Fe^{2+} in hemoglobin and myoglobin (and wins).

Cytochromes are a class of enzymes that catalyze the oxidation of carbohydrates, an essential energy reaction in the body.

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) + \text{Energy}$$

The cytochrome enzyme molecule aids this reaction by being a repository for (and source of) electrons.

$$\text{Enzyme-protein-}(Fe^{2+}) \rightleftharpoons \text{Enzyme-protein-}(Fe^{3+}) + e^{-}$$

As with hemoglobin, the iron ion forms an octahedral complex. Unlike hemoglobin, the iron ion in cytochrome has all six positions firmly bonded; thus iron's only action is to accept and then give off an electron.

A copper compound, hemocyanin, serves the same oxygen-carrying function in lobsters and oysters that hemoglobin serves in higher animals.

Earth Science

About three billion years ago primitive plants began to produce oxygen as a product of photosynthesis. One intriguing observation is the slow rate at which oxygen accumulated in the primitive atmosphere. The element iron in earth's crust is the key. Prior to the existence of free oxygen, molten iron combined with silica to form iron silicates. When free oxygen became available as the result of photosynthesis, iron combined with it to form abundant quantities of iron oxides. So the earth rusted, using up free oxygen until all elementary iron near the surface was combined. Then the free oxygen accumulated in the atmosphere and ultimately allowed for the appearance of land-based animals.

TO THE CONTEMPORARY WORLD

Personal

1. Transition elements are of considerable commercial importance. The metallurgy industry employs many chemists.

2. A knowledge of the properties of metals such as copper, silver and gold is advantageous to jewelers.

Community

Field Trips

- To a plating plant.
- To a film developing business to observe the role of silver in this process.

Knowledgeable Individuals

- Dentist—use of mercury amalgams for filling teeth.
- Environmental scientist—role of heavy metals in environmental pollution.
- Jeweler—use of gold and silver.
- Artists who work with metals.
- Physician—role of iron in human health.

1. **Color in gemstones**. The color in most gemstones comes from transition metal ions present as impurities in a parent mineral. For example, rubies are simply the mineral corundum, Al_2O_3, with Cr^{3+} ions replacing a few Al^{3+} ions in the structure. Blue sapphires involve Fe^{2+} and Ti^{4+} ions replacing some Al^{3+}. Emeralds are the mineral beryl, $Be_3Al_2(SiO_3)_6$, with Cr^{3+} ions replacing a few Al^{3+} ions (see *The Chemistry of Rocks, Minerals and Gems* module).
2. **Bluing**. Prussian Blue (also called Paris blue, Berlin blue, Hamburg blue, and Chinese blue) is the salt $Fe_4[Fe(CN)_6]_3$, formed from Fe^{3+} cations and the iron(II) complex $Fe(CN)_6{}^{4-}$. Prussian Blue is the bluing agent in laundry products and the source of color in engineering blueprints.
3. **Bronze statues**. Bronze is an alloy containing 90% copper and 10% tin. Bronze statues form a green surface coating as they age, which actually protects them from further weathering. The reaction involves copper and constituents of moist air:

   $$2Cu(s) + H_2O(l) + CO_2(g) + O_2(g) \rightarrow Cu_2(OH)_2CO_3(s)$$

   The green product can be viewed as a solid-state mixture of $Cu(OH)_2$ and $CuCO_3$.
4. Many transition elements contribute to the colors in major artists' pigments. Some examples are chrome yellow ($PbCrO_4$), manganese violet [$(NH_4)_2Mn_2(P_2O_7)_2$], the iron(III) oxide reds, which go by many names including Venetian red, Turkey red and Indian reds (Fe_2O_3), cobalt blue [$CoO{\cdot}Al_2O_3$], blue copper phthalocyanine ($C_{23}H_{16}N_8Cu$) and molybdate orange (mixture of $PbCrO_4$, $PbSO_4$, and $PbMoO_4$).
5. Transition elements are commonly used in industry and medicine, for example:
   a. Titanium metal is used in place of iron in both aircraft and racing bicycles. Titanium tetrachloride ($TiCl_4$) is a volatile liquid whose vapors react with air to give a dense smoke of TiO_2. This reaction was once used by the U.S. Navy to make smoke screens during naval battles. Titanium dioxide (TiO_2) is the most common white pigment in paint. It is also used as a brightener and a sunscreen.

b. Chromium is mixed with nickel and iron to produce stainless steel. Nichrome, an alloy of chromium and nickel, is often used as the wire heating element in toasters.

c. Cobalt is used to prepare an alloy called alnico, which forms powerful magnets. Gamma radiations from cobalt-60 are used in cancer treatment.

d. Cadmium is used to make nickel-cadmium batteries.

e. Mercury dissolves many other metals to form solutions called amalgams. Dentists use a silver amalgam to fill teeth. This amalgam contains an excess of silver. Mercury(II) nitrate [$Hg(NO_3)_2$] was once used in the manufacture of felt for hats. Workers often developed severe mercury poisoning. This affliction leads to nervous disorders, loss of hair and teeth, memory loss, and insanity. The phrase "mad as a hatter" refers to these symptoms of mercury poisoning.

f. At one time platinum was known as "white gold," but this term is now applied to an alloy of gold and platinum. The most exciting use of platinum compounds is as the cancer chemotherapy agent cisplatin [$(H_3N)_2PtCl_2$]—see *Chemistry in Medicine* module. Finely divided platinum, functioning as a catalyst, is used to reform straight-chain hydrocarbons with low octane numbers into their branched-chain isomers, which have higher octane numbers.

Extensions

1. If a spectrophotometer is available, students may enjoy analyzing colored solutions. (Laboratory Assesment Builds Success, Activity 18)
2. Many chromium compounds are alleged carcinogens, yet there is evidence that chromium is necessary to health. This would make a good library research project.
3. Chlorophyll and hemoglobin are very similar. Chlorophyll contains Mg, while hemoglobin contains iron. Hemocyanin, found in certain marine animals (such as squid), contains a heme-like porphyrin with a Cu^{2+} ion in place of Fe^{2+}. Invite students to investigate these substances, their formulas and properties.
4. In some localities small quantities of transition metals are allowed to enter the water supply through discharge into streams and lakes. An interesting Science Fair Project would be to find a method to determine trace quantities of hazardous ions in your community's streams.
5. The history of detecting mercury in fish is interesting because it depends in part on increasingly sensitive detection methods. Students could compare methods used to determine mercury 50 years ago to methods used today. The *World of Chemistry Secondary School Project* has an introduction to such a project: Tape 2, Measurement.
6. Natural deposits of transition element ores are not evenly distributed on earth. Have each student choose one transition element and locate where its ores are concentrated. Plot the sites on a world map and speculate about economic and political consequences.

References

Module developed by Diana Doepken-Rischling, Michael Pavelich, and Patricia Smith, the Colorado (South) team.

Dreyfus, C., and Dreyfus, H. (1984). *A one pot reaction,Volume III*. A Camille and Henry Dreyfus Institute Curriculum Module.

Emsley, J. (1991). *The Elements* (2nd Ed.). Oxford: Clarendon Press.

Key facts and data about the chemical elements that make up the world around us are brought together in a systematic and easy to use format.

Gardner, M., Heikkinen, H., and Smith, P. (1990). *Laboratory assessment builds success*. Madison, WI: Institute for Chemical Education.

Eighteen laboratory activities for secondary level that feature various patterns of laboratory activity including computer interfacing, microscale techniques, and a variety of hands-on, minds-on efforts to make important chemistry concepts understandable to high school students.

Greenwood, N.N., and Earnshaw, A. (1984). *Chemistry of the elements*. New York, NY: Pergamon.

Presents a balanced, coherent and comprehensive account of the chemistry of the elements for both undergraduate and post-graduate students. Full of ingenious experiments, intriguing compounds and exciting new discoveries.

Kolb, D. (1988, November). *Journal of Chemical Education, 65,* 1004.

A colorful demonstration showing six oxidation states of manganese on the lighted stage of an overhead projector.

Kolb, D. (1987, April). *Journal of Chemical Education, 64,* 349.

Some demonstrations illustrating basic chemical concepts of transition metals or complex ions that can be carried out on the lighted stage of an overhead projector.

Russo, T. (1990). *Microchemistry 2*. WestChester, OH: Kemtec Educational Corp., 9889 Cresent Park Drive.

This microchemistry laboratory manual contains an experiment about the transition metal vanadium.

Appendix

- **Transparency Master**
  1. Word Search
- **Humor**

Word Search

D	N	A	G	I	L	O	X	M	X	O	B	X	L	K	N	S
B	F	O	D	H	Y	C	H	N	P	U	R	P	L	E	O	J
I	W	U	L	U	G	U	H	D	T	L	V	D	M	S	I	P
Y	N	M	I	W	O	N	D	R	A	O	Z	T	C	T	T	L
N	R	Z	O	P	N	R	O	R	O	D	O	D	Q	U	A	H
L	E	P	N	Y	G	R	D	I	N	M	W	N	H	M	N	U
E	P	N	I	U	D	E	W	M	T	L	I	F	R	G	I	Z
W	P	S	Z	L	H	N	U	W	T	I	E	U	X	Z	D	K
K	O	I	A	A	U	R	X	D	Y	M	S	W	M	I	R	I
I	C	V	T	K	E	E	A	T	J	C	L	N	D	X	O	G
G	N	C	I	N	D	W	R	R	S	D	E	S	A	H	O	Y
F	O	N	O	V	M	K	C	V	M	K	S	U	C	R	C	K
M	V	R	N	R	A	J	N	O	R	I	Z	C	G	H	T	C

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1. Molecule or ion surrounding a metal ion in a complex.
2. Adjective for the *d*- and *f*-block elements.
3. Shape associated with six species equally spaced around a central species.
4. Transition element identified by electron configuration [Ar]$3d^6 4s^2$.
5. Color of aqueous solution of permanganate ion.
6. In addition to tin, bronze statues contain this transition metal that turns them green upon aging.
7. Transition element responsible for the red color in rubies.
8. Type of energy required to remove an electron from an atom.
9. Type of complex composed of a metallic ion and its surrounding ligands.
10. Swiss chemist who predicted geometries of complex ions.

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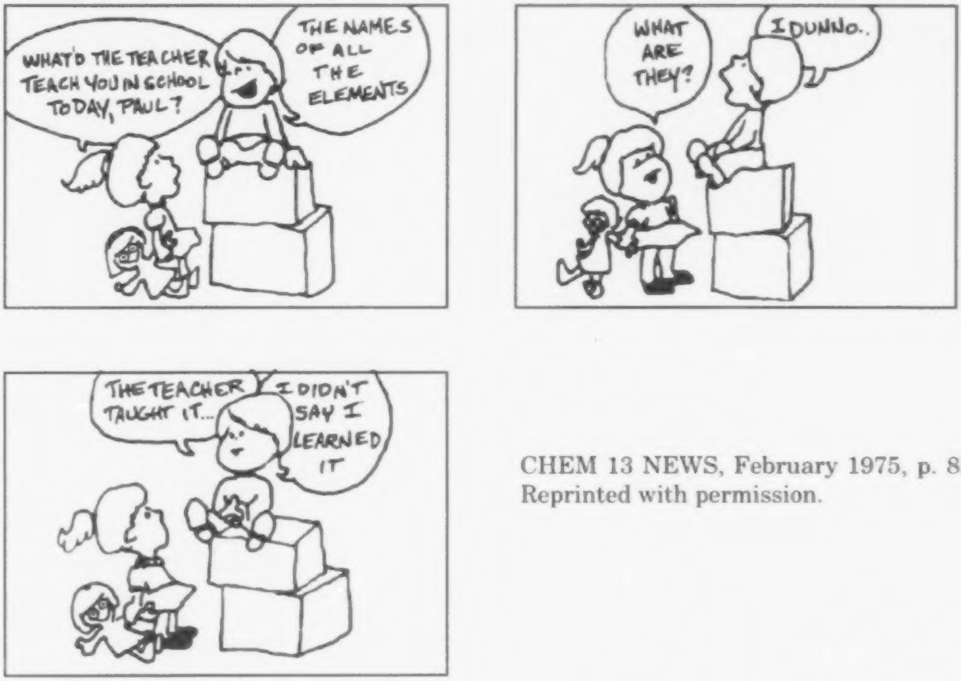
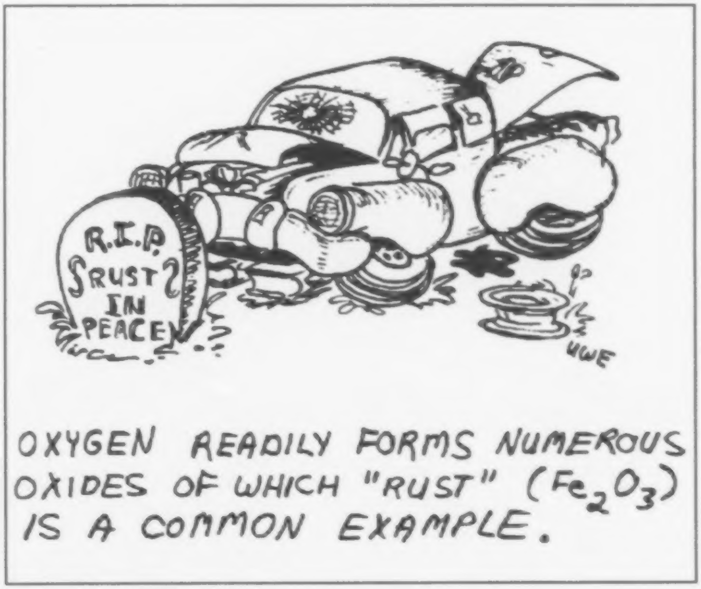

WHAT'D THE TEACHER TEACH YOU IN SCHOOL TODAY, PAUL?
THE NAMES OF ALL THE ELEMENTS
WHAT ARE THEY?
I DUNNO..
THE TEACHER TAUGHT IT...
I DIDN'T SAY I LEARNED IT

CHEM 13 NEWS, February 1975, p. 851. Reprinted with permission.

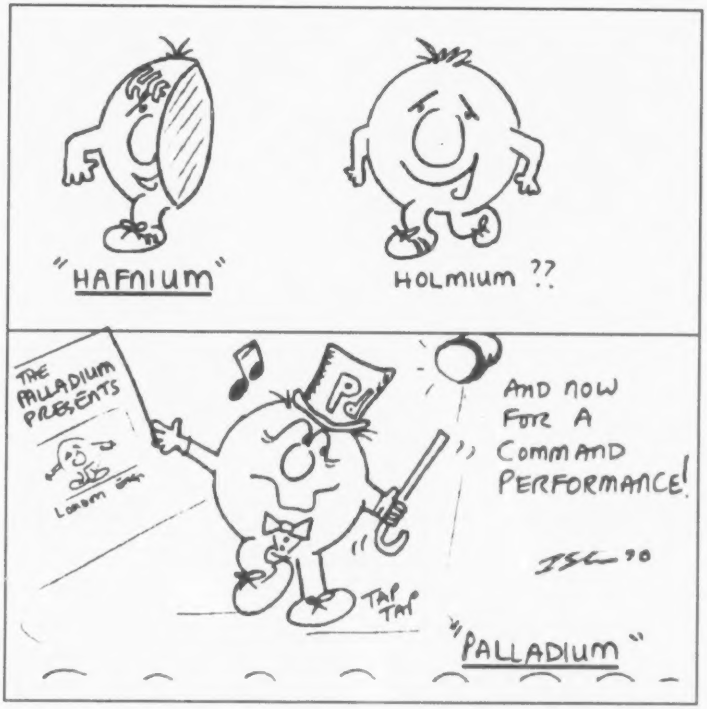

R.I.P. RUST IN PEACE
UWE
OXYGEN READILY FORMS NUMEROUS OXIDES OF WHICH "RUST" (Fe_2O_3) IS A COMMON EXAMPLE.

CHEM 13 NEWS, December 1980, p. 8. Reprinted with permission.

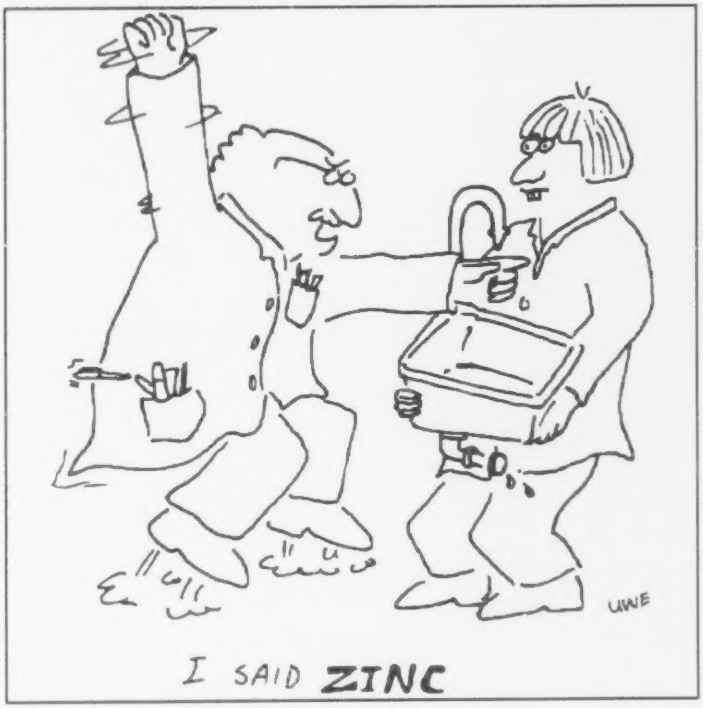
Fe
Zn
Ag
Pb
Au
Hg
Hg
Fe
Zn
Mn
Mn
Sn
Pb
Ag
Pb
Sn
Au
Sn
Ag
Pb
Hg
Zn
Fe

A heavy metal concert!

CHEM 13 NEWS, December 1983, p. 3. Reprinted with permission.


"HAFNIUM"
HOLMIUM ??
THE PALLADIUM PRESENTS
Pd
AND NOW FOR A COMMAND PERFORMANCE!
TAP TAP
JSL '90
"PALLADIUM"

CHEM 13 NEWS, September 1991, p. 6. Reprinted with permission.


UWE
I SAID ZINC

CHEM 13 NEWS, March 1979, p. 5. Reprinted with permission.

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A SourceBook Module

Version 2.0 1994

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

COMPUTER USES

Computer-Assisted Instruction

As *ChemSource* is released for teachers and begins to be used by them, software tools will be available providing computer-based solutions for many skills traditionally targeted as learning goals in conventional courses—stoichiometry, balancing equations, electrochemical cells, electron dot structures, *etc*.

Teachers of traditional disciplines (chemistry, mathematics, English, art, music) are placed in difficult positions. Clearly, computers are changing what content it is important to know. But, since this transition is just beginning, one has no experience upon which to base instruction.

This module stresses computer tool use. Few, if any, schools will have computers for each student in each class. Most schools have several computer laboratories, and the typical chemistry laboratory classroom has one or more personal computers.

Computer use in chemistry classes can be divided into four broad groups:

- Computer-assisted instruction
- Use of computer applications for general productivity
- Use of computer applications for profession-specific productivity
- Use of computers in laboratory experiments

During the period 1980-90, many drill and practice programs were developed for a wide variety of personal computer platforms that afford students opportunity to employ common chemistry applications skills. This is a dynamic area, especially with new and more powerful computers becoming increasingly available in schools. The introduction of multimedia programs in 1984 and hypermedia programs in 1987 promises to revolutionize this effort still further—with teacher-created software becoming commonplace.

This module does *not* deal with computer-assisted instruction. Hypermedia materials have made it rather simple to make small-sized computer databases interactive with videodisc players. Thus the *Periodic Table* and *Redox* videodiscs available from Project SERAPHIM make new and exciting types of instruction possible for teacher presentation, use by individual students, and cooperative learning.

There is a variety of software available for computer-assisted instruction in chemistry, ranging from simple drill and practice exercises to complex multimedia simulations. Software for direct student instruction should be carefully reviewed. Poorly written software can frustrate and discourage students, whereas good instructional software can give students learning experiences not available by other means. For example, students can perform simulated experiments on a computer using equipment or reagents not ordinarily available. They can also repeat the experiment many times, changing conditions to see what happens.

Good tutorial and drill-and-practice software is infinitely patient, allowing each student to proceed at his or her own speed while giving each student feedback specific to their answers. Good instructional software can enhance student learning and adjust to individual student needs. Under the heading of *Media*, all of the *SourceBook* modules list computer materials that might be useful in specific contexts.

Computer Applications for General Productivity

Essentially all professional, white collar, and clerical workers as well as many blue collar workers use computers. Generic application programs provide the best introduction to computing. Four types of programs most useful to chemistry teachers are:

- Word processing
- Spreadsheet (including graphing)
- Graphics (paint, draw)
- Communications

WORD PROCESSING (WP)

Foremost among your students' applications of computers will be word processing. Laboratory reports and other written assignments can be turned in on floppy discs, perhaps with accompanying hard copy. An even more appropriate strategy is to use some form of electronic mail between student and teacher. The biggest factor limiting extensive use of word processing is student access to machines, a problem that is diminishing.

Given hardware and software, however, you can try new strategies:

Group Writing

When students work in groups, writing the report sections can be assigned to group members. Each group member would use the word processor to prepare the assigned section. One group member would be responsible for bringing all sections together. Software companies continue to facilitate group uses of word processing on the same document—many users accessing the same document, often simultaneously.

Editing

Require that all reports be edited. Do not accept reports as final until appropriate editorial corrections have been made. The skills of editing reports are evolving. Encourage the use of spell checkers, grammar checkers, and thesaurus programs.

Other Writing

Assignments such as having students write letters to the editor of local newspapers regarding environmental impacts of chemicals in the community can be included as a word processing activity.

SPREADSHEETS

Because both chemistry teachers and spreadsheet programs deal with numbers, a symbiosis between the two seems natural. Actually, spreadsheets offer opportunities that have only begun to be explored.

Two kinds of activities are generally conducted in the teaching laboratory. Activities where every student in class performs an experiment to get numerical data are very common. For example, everyone might be trying to determine the formula of a hydrated salt (*e.g.*, find x in $CuSO_4 \cdot xH_2O$). In this case, the spreadsheet should be used to compare the results of all the students in the class. The teacher or one designated student would supervise this computer use, and all students contribute their results—either final results, or raw data.

In the second type of activity, every student gathers a set of data and tries to analyze that data to draw conclusions. For example, everyone might be measuring the volume of air trapped in a sealed syringe based upon the weight pushing down on the syringe's plunger. In this case, each student will want to use a spreadsheet to help analyze the set of data.

Both kinds of data analysis are important, but neither received much attention prior to 1990 because of the tedium of performing repeated calculations. The spreadsheet minimizes this tedium.

Modern spreadsheet programs have built in or have companion graphing programs. These will plot data in various ways, and even find best-fit equations for sets of data. If you use a spreadsheet that does not have the capability of creating graphs from data, particularly the *xy* graphs used so often in science, then companion graphing programs must be made available.

The analyses and graphs prepared using a spreadsheet can be imported into the word processing program for inclusion in the students' reports.

Many computations can be made automatic by including formulas for them in a spreadsheet. Consider the following tables shown in Figure 1 for compounds of sodium, sulfur, and oxygen. These were obtained from one spreadsheet with the first column holding atomic symbols, the second atomic masses, the third an arbitrary, student-entered number of atoms per unit, and the fourth the percentage by mass of that element in a compound of the formula shown above. The entry at the lower right is the formula mass.

			Na_2SO_3
Na	23.0	2	36.5%
S	32.0	1	25.4%
O	16.0	3	38.1%
			126.0

			Na_2SO_4
Na	23.0	2	32.4%
S	32.0	1	22.5%
O	16.0	4	45.1%
			142.0

			$Na_2S_2O_3$
Na	23.0	2	29.1%
S	32.0	2	40.5%
O	16.0	3	30.4%
			158.0

Figure 1. Percent by mass composition: A typical spreadsheet.

These computations are done using only two formulas. The formula in the cell at the bottom right is =B2*C2 +B3*C3 + B4*C4. (Columns are labeled with letters; rows are labeled with numbers. This small array has A1 as the cell at the top left and D5 as the cell in the bottom right.) This formula calculates the formula mass of the substance.

The formula in cell D2 just beneath the chemical formula is =IF($D$5 =0,0,100*B2*C2/$D$5). (The dollar signs in the formula are used to identify specific cell addresses that do not change; this is important when copying and pasting formulas.) This formula checks the number in the lower right (*i.e.*, the formula mass). If it = 0, then it puts the number 0. If not, then it puts the product of 100 times the number of atoms times the atomic mass divided by the formula mass. Essentially the same formula is copied into cells D3 and D4 as =IF($D$5 =0,0,100*B3*C3/$D$5) and =IF($D$5 =0,0,100*B4*C4/$D$5).

The specific syntax and functions employed will vary from spreadsheet to spreadsheet.

The theory part of a course can also make use of spreadsheets. For example, during one editing cycle of the *ChemSource* materials the editors engaged in a discussion about weak electrolytes. There is a confusing aspect about weak electrolytes. On the one hand, the extent of dissociation of weak electrolytes increases with decreasing concentration. On the other hand, the conductivity decreases with decreasing concentration. The following arguments can be made. For a weak acid such as acetic acid (represented as HOAc), the dissociation equation is:

$$\mathrm{HOAc} \rightleftharpoons \mathrm{H^+} + \mathrm{OAc^-}$$

For this reaction, an equilibrium expression is written: $K_a = \dfrac{[\mathrm{H^+}][\mathrm{OAc^-}]}{[\mathrm{HOAc}]}$

The conductivity of the solution will be related to the sum of the ion concentrations:

$$\text{Conductivity } \alpha \ [\mathrm{H^+}] + [\mathrm{OAc^-}]$$

Suppose that the initial concentration of HOAc is represented by the symbol C. Suppose the concentration of HOAc that dissociates at equilibrium is represented by x. Then, to a first approximation, at equilibrium $[\mathrm{H^+}] = [\mathrm{OAc^-}] = x$ and $[\mathrm{HOAc}] = C - x$. At equilibrium, $K_a = x^2/(C-x)$. Also, the percentage dissociated is given by $100\%\ (x/C)$. Finally, the conductivity is directly proportional to x. Let us rewrite the equilibrium expression as given in Equation 1.

$$x^2 + Kx - KC = 0 \tag{1}$$

Using the formula for the roots of a quadratic equation $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

and solving for the appropriate root (non-negative concentration) with $a = 1$, $b = K$, and $c = -KC$ gives Equation 2:

$$x = \frac{-K + \sqrt{K^2 - 4KC}}{2} \tag{2}$$

This equation can be rewritten using appropriate notation into a "cell" of a spreadsheet, solved for a particular value of C (such as 1 M), and then copied so that the same formula is used to calculate the quantities of interest for other concentrations. The following calculations were performed in this manner:

From these data, the following graph was plotted. In this graph, the circles ○ represent the concentration of the ions resulting from dissociation and the squares □ represent the percent dissociation. The circles track the conductivity, which decreases as the concentration is lowered, while the squares track the amount of dissociation, which increases as concentration lowers.

$K = 1.8 \times 10^{-5}$

C	x	100x/C%
1.00	0.0042	0.42%
0.90	0.0040	0.45%
0.80	0.0038	0.47%
0.70	0.0035	0.51%
0.60	0.0033	0.55%
0.50	0.0030	0.60%
0.40	0.0027	0.67%
0.30	0.0023	0.77%
0.20	+0.0019	0.94%
0.10	0.0013	1.3 %

Figure 2. Calculation of (100x/C)% using decreasing values of C.

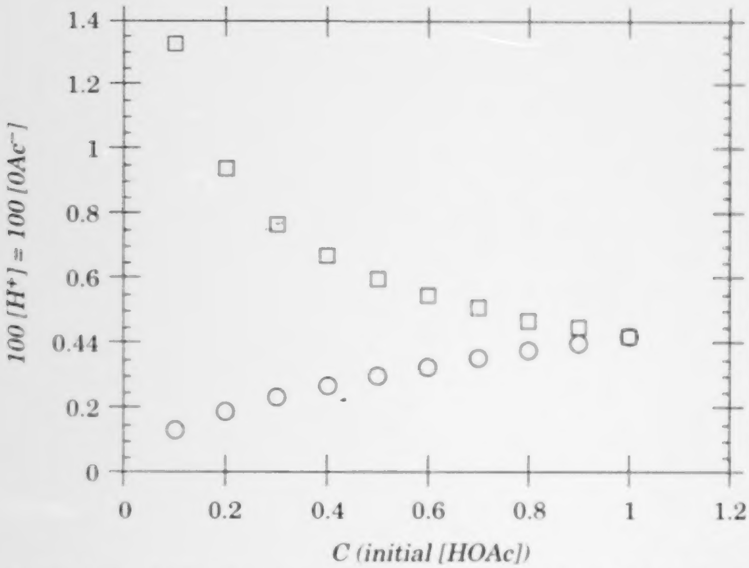


Figure 3. Plot of data from Figure 2. **(□ 100 x/C = % dissociation; ○ 100 [H+] =100 [OAc–]**

As the graph indicates, x, the concentration of H^+ and OAc^- increases as the concentration of HOAc increases. This means that the conductivity increases with increasing initial con-centration. However, at the same time, the percentage of molecules in this electrolyte that are dissociated into ions decreases with increasing concentration. Until students play with these numbers to see how they balance out, there is an apparent discrepancy or contradiction. The following spreadsheet cell equations (Figure 4) were used to create the data table of Figure 2.

K = 0.000018

C	100x	100x/C %
1.0	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A4*$B$1)))/(2)	=B4/A4
0.9	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A5*$B$1)))/(2)	=B5/A5
0.8	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A6*$B$1)))/(2)	=B6/A6
0.7	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A7*$B$1)))/(2)	=B7/A7
0.6	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A8*$B$1)))/(2)	=B8/A8
0.5	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A9*$B$1)))/(2)	=B9/A9
0.4	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A10*$B$1)))/(2)	=B10/A10
0.3	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A11*$B$1)))/(2)	=B11/A11
0.2	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A12*$B$1)))/(2)	=B12/A12
0.1	=((–1*$B$1) + SQRT(($B$1*$B$1) + 4*(A13*$B$1)))/(2)	=B13/A13

Figure 4. Spreadsheet cell equations used to create data in Figure 2.

GRAPHICS PROGRAMS

Graphics programs provide pictures. *Paint* programs create images by assigning a color value to each point on a computer screen (or printer page.) They may assign just black or white, but shades of gray and colors are commonplace. *Draw* programs create "computer" objects defined by equations and algorithms. The image appears because the computer draws the image to the specifications of an equation rather than recall and assign a color value to each spot (pixel) on the screen.

If a picture is worth one thousand words, then students should be creating pictures. Painting/drawing programs may be used to develop sketches of laboratory hardware, draw molecular structures, depict chemical reactions, *etc*.

Importing digitized pictures of various resources such as laboratory equipment is possible. These pictures can be modified using appropriate software packages.

COMMUNICATIONS

The National Science Foundation-sponsored worldwide communications network called Internet includes many K-12 teachers. By April 1993, over 2500 teachers in Nebraska had Internet addresses, for example. Internet provides many communication links, the simplest of which is electronic mail. Communication requires software and a modem or some other software/hardware link to a computer that is part of the Internet.

Connections to Internet aside, teachers and students can communicate with one another when both can access computers with modems. This sort of connection often employs intermediaries called electronic bulletin boards, some of which may be accessed by 800-numbers. Telephone communications costs usually prohibit extensive use of this form of communication.

One of the features provided by Internet is the ability to log onto list servers. A chemistry teacher list server allows teachers to post questions and comments in such a fashion that hundreds of persons can participate vicariously while a few exchange comments.

Electronic communication or E-mail over the Internet is a straightforward matter. To send E-mail to someone, the sender needs only to know the intended receiver's E-mail address. To reach one of the authors of this module, for example, send E-mail to dbrooks@unlinfo.unl.edu. Communication to the National Science Foundation is a straightforward matter If one wanted to communicate with some Foundation representative, let's just say someone named Donald Jones, the first guess at the electronic mail address would be djones@nsf.gov. Of course, if the Foundation hired someone named Donna Jones, then that person would need to be assigned a different address. Address conflicts are inevitable. As of this writing, few Internet sites have implemented ways of looking up E-mail addresses of individuals connected through those sites. Getting someone's E-mail address can be as frustrating as getting their phone number — even more so, since calling them may be the most reasonable access to their E-mail address. In mid-1993, the future of Internet was clouded as attempts were being considered to privatize this system.

Many school districts already link the administrative headquarters of their various buildings using E-mail. In a growing number of school systems, all teachers within a building are linked together by computer-based E-mail. These networks are called local area networks, or LANs. Some companies have implemented still wider networks, WANs.

Besides direct communication between students, how might teachers include communications activities within instruction? The simplest versions might involve an experiment worked on by two chemistry classes in the same building in different periods where students in the early period start an experiment that is completed by students attending a session later that day. Communications between these student groups might use a single in-classroom computer as their link.

Cooperative experiments performed by students at different sites already constitute a part of some early network experiments such as *KidNet*. It is likely that a variety of national computer experiments will emerge in the mid-1990s. These are likely to include both environmental and ecological studies.

Professional chemists have used communications programs for some time to connect to large databases. Most of the database programs that chemists use reside on large mainframe computers. In addition to the hardware and software, a teacher needs some means of covering costs for accessing the database as well as covering phone line charges. Searching databases such as *Chemical Abstracts* can be accomplished using so-called on-line computer searches.

Computer Applications for Profession-Specific Productivity

Some powerful worksheets for spreadsheet programs have been developed. The idea here is that the chemical know-how is preprogrammed into a so-called template. The user can use standard spreadsheet software together with the template to obtain complex results without the need for programming the spreadsheet. They run the template, but need to enter only the specific parameters for their problem—the formula and mass of a reactant; the name of a gas with its pressure and temperature, *etc.* For chemistry templates, nearly all of the needed chemistry know-how is built in.

The *ChemSource* beta CD-ROMs developed and distributed in 1991 and 1992 included many teacher-specific chemistry tools. One of these tools, Chemistry Laboratory Assistant, affords teachers the opportunity to keep close control of managing experiments. It includes the ability to scale amounts in procedures for preparing solutions to fit the number of students/sections, manage inventory, and create chemical purchase orders. These features are tied to specific experiments. The CD-ROM version of *ChemSource* includes similar support for all of the included experiments and demonstrations.

Computer applications are being developed with specific professional needs and applications in mind. In chemistry, programs that display 3-dimensional molecular structures in 3-dimensions are at the forefront. In fact, some of these systems are so powerful that they compute structures (bond lengths, bond angles) that are more reliable than some experimentally determined structures. Very few high schools have computers powerful enough to run these programs.

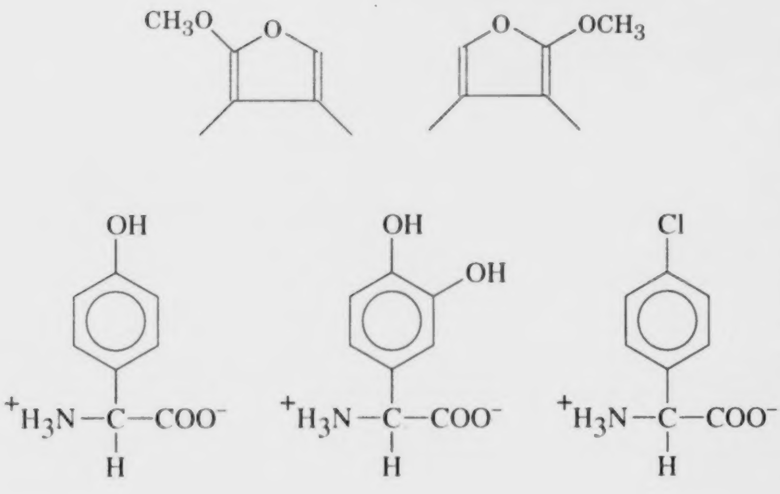


Figure 5. Structures drawn using a molecular drawing program.

Although expensive, molecular draw-ing programs are within the reach of many secondary schools. These run on ordinary computers. The structures shown in Figure 5 were drawn quickly using a molecular drawing program.

Access to chemistry specific software is likely to change drastically over the upcoming decade. When *ChemSource* is first published, a series of very small portable computers called "pro-fessional assistants" or some other name will emerge. Several hand-held calculators already have been produced that aid in chemical calculations. The power and pro-grammability of such devices will be enhanced enormously, and the use of these tools in laboratory work is likely to become commonplace.

Computer Interfacing with Experiments

The first major uses of computers in laboratories were in the area of assisting scientists to obtain data. Computers were used to record signals generated by probes (time, temperature, pressure, pH, *etc.*) and sometimes to control experiments on the basis of these signals. Data analysis followed shortly thereafter.

Significant initial uses of computers by high school chemistry teachers involved interfacing. Using the game ports of Apple II and IBM personal computers, teachers constructed and attached temperature probes and colorimeters. Ten of these experiments appear on the F-disc of the *Doing Chemistry* series.

In order to "interface" an experiment, a probe, an interface and some software are needed. In the simplest cases, a game port serves as the interface. The interfaces perform several functions including timing data acquisition and converting analog signals of the probe into digital signals suitable for the computer. In an 8 bit computer, only 256 values of the digitized signal are possible, and this limits experimental precision.

Strategies for Teachers

PROVIDING ACCESS

Many students have access to computers in their homes. It is a good idea to survey them early in the school year to get an idea of just what hardware and software are available outside of the classroom.

INDIVIDUAL COMPUTERS IN LABORATORY CLASSROOMS

When one or two computers are available in the classroom, these can be set up as learning centers. While students work in laboratory, the computer should always be in use. When experiments are performed that do not make use of computer controlled sensors, students still should have access to the computer for scientific writing, *etc*.

For example, asking students to prepare a library paper or a marketing plan, a biography about a scientist, or even a "biography" about a chemical element of their choice is often used as a chemistry class assignment. If this assignment is made several weeks into the course but collected at the end of a long instructional unit (such as near the end of a semester), then each student can be given 1-period-long access to the class computer to work on their paper during class. This is a way to help focus students on the use of computers as scientists' tools even if the only work done is word processing.

A teacher can afford students even more opportunities for creativity. Using hypermedia and videodiscs, it is possible for students to prepare electronic materials for assignments. Students might write a hypermedia description of an element or family of elements in which they design access to images through databases such as the *Periodic Table Videodisc* (*JCE: Software*, Vol. I, No.1). The software for this sort of creative work by students already is in place; such assignments already have been included in elementary school curricula.

CD-ROM technologies also are becoming available. One principal means for dissemination of *ChemSource* is CD-ROM. Soon students will be able to access very powerful databases in this medium. Again, opportunities for teachers to provide creative opportunities for students abound within the context of this hardware/software environment.

When "data" laboratories are assigned, one student can be given the task of entering all the data into a spreadsheet program. Students can see the same computer screen when a liquid crystal display or some similar device is available. (A TV camera system can be used to project an image of a computer screen through a television. Although some careful camera work may be required to keep the clarity of the projected image adequate, this simple kind of through-the-air projection system is sometimes easier for classroom teachers to utilize than one requiring specialized overhead display equipment.)

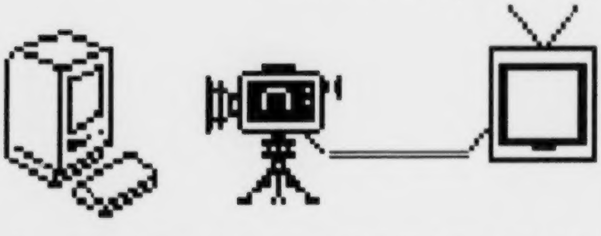


Figure 6. Computer screen to a TV camera to a TV screen.

Spreadsheet data should be analyzed whenever possible. For example, suppose there is an experiment where 12 different-sized objects are available, 6 made of aluminum and 6 of magnesium. A class of 24 students, working in pairs, can mass these objects. Divide the class into 12 student pairs, and give each pair one object. The students mass the objects hanging from a hook on a triple-beam balance, once in air, and once submerged under water. All of the data may then be analyzed using a "collective" spreadsheet. The data—two data points per student pair—are entered, analyzed, and then discussed.

Another way to tackle the same activity is to choose 12 different-sized objects of the same material and assign one each to a student pair. In this case, have six student pairs measure the mass in air and when submerged in water. Have the other pairs measure the mass in air and the volume by displacement of water in a cylinder. This data may then be spreadsheet analyzed. The expected result is that the error associated with the volume-by-measured-displacement method will be greater than that observed when two masses are measured. The shared spreadsheet analysis will make any differences very apparent. Be certain to lead a class discussion about these differences.

COMPUTER LABORATORIES

Computer laboratories are not desirable as means of providing computer access. In contemporary workplace situations, workers use computers at their workstations, not in some remote room. The reality of funding makes it likely that schools will use computer rooms for some time. When this is the case, be sure to provide excellent software that permits easy exchange of materials between students at different terminals and between programs at one terminal. This implies using MultiFinder or System 7 at Macintosh stations, or Microsoft Windows 3.1 or higher at DOS stations.

USING TEAMS

As computers change the meaning of the term "professional," they also change the ways in which professionals function. Today professionals often work in teams. Any experiments students perform working in teams is an appropriate reflection of contemporary practice, one expected to persist into the next century. Report writing is one area in which teams can be used. For example, ask each three-person group to assign one of its members to:

- Writing using a word processor;
- Data analysis using a spreadsheet; and
- Visual presentation using a painting/drawing or graphing program

When multitasking, computer environments such as Macintosh MultiFinder or System 7, or MS/DOS Windows are available, the merging of information into one report is a fairly straightforward matter. (In actual practice, much of this sharing would be accomplished electronically, with team members electronically mailing their results around their building, or even around the world.)

Consider the small-scale Boyle's Law experiment in which a gas sample trapped in a sealed insulin syringe is compressed using increasing masses (forces) applied to the plunger by adding water to an open vessel placed on top of the syringe (see *Doing Chemistry*, Experiment D04). The data shown in Figure 7 are typical of those obtained during the experiment.

Water Added, mL	Volume, mL
0	29.5
500	25.1
1000	22.5
1500	20.0
2000	18.2
2500	16.7
3000	15.5
3500	14.2
4000	13.4

Figure 7. Boyle's Law data.

A spreadsheet may be used to analyze these data, and to generate the relationship between pressure and the volume (Figure 8) and the reciprocal volume (Figure 9).

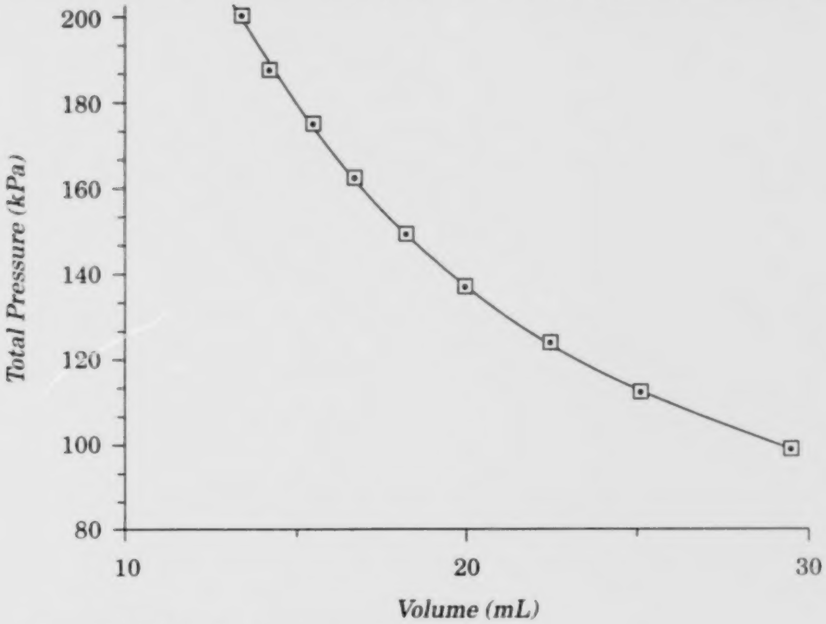


Figure 8. Plot of pressure vs. volume.

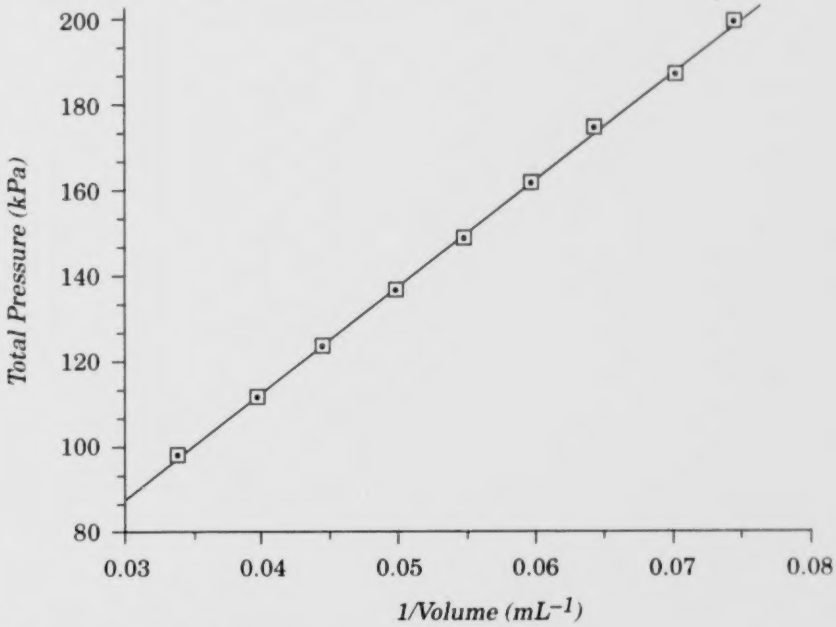


Figure 9. Plot of pressure vs. reciprocal volume.

Also, a picture of the apparatus can be prepared using a painting program (Figure 10). This particular image was obtained using a digitized video picture taken directly from the laboratory apparatus. Inexpensive image digitizing equipment and software are available for modern computers, and can use the images from TV cameras or camcorders found in schools as the source of the images to be digitized.

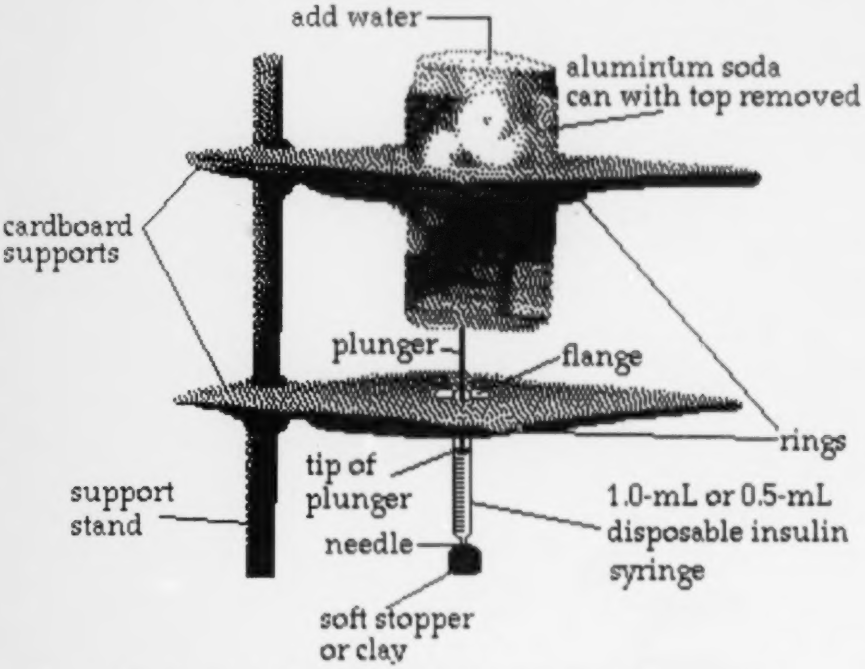

add water
aluminum soda can with top removed
cardboard supports
plunger
flange
rings
tip of plunger
support stand
1.0-mL or 0.5-mL disposable insulin syringe
needle
soft stopper or clay

Figure 10. Picture of apparatus done with a paint program.

Another way to use images from textbooks, magazines, and other resources as sources for electronic files is through scanners.

As a teacher, you might want to be certain that each student gets to play each of the different team roles—the writer, the analyzer, and the graphist—at least once. However, permitting students to become specialists and to emphasize one of these roles as an area of personal expertise is a better reflection of contemporary practice than is insisting that all students rotate uniformly through all "specialties."

Teachers can try to engage students in completely different ways. For example, one class can design experiments and transmit those specifications via E-mail to a second class that conducts those experiments and provides results that, in turn, are analyzed by the first class. Providing this sort of experience with a pair of classes and two experiments may really give them a flavor of both roles in science—group leader and technician—typical of many industrial and government settings.

Suggestions for ChemSource Experiments and Demonstrations

Several of the *ChemSource* experiments lend themselves to specific types of extensions using computers.

Atomic Structure

Identification of Substances through Atomic Emission of Light
Set up a spreadsheet so that, when a number is entered for a wavelength in one cell, the frequency and energy are computed and presented in adjacent cells.

Atomic Structure

Locating an "s" Electron in an Atom by Analogy
Number each point on the graph paper. Measure the distance between the point and the center bull's eye. Use this data on a spreadsheet for analysis.

Chemical Bonding

Metal and Ionic Crystals
Use a CAD program to create 3-D models of spheres packed together.

Chemical Bonding

Displaying Structures in Two Dimensions
Use a molecular structure drawing program to draw simple structures.

Condensed States

Density, A Straight Line Function
Analyze the data using a spreadsheet.

Condensed States

Energy Needed to Melt Ice
Gather data from all students and analyze using a spreadsheet. Compute a value for each set of data. Then analyze the results collectively.

Condensed States

Some Like It Hot
Analyze the data using a spreadsheet.

Gases

Volume and Temperature Changes in a Gas
Pool the data from working groups and analyze the data for the entire class. Use a graphing application or a spreadsheet graphing feature.

The Mole

Understanding the Mole
Pool the data from working groups and analyze the data for the entire class using a spreadsheet.

The Mole

Percent Composition of Magnesium Oxide
Analyze the data using a spreadsheet. Collect the data and analyze each set separately. Then combine the results for the entire class.

Molecular Geometry

Building Molecular Models
Use a molecular structure program that draws 3-D representations of molecules and compare these structures to the models constructed.

Rates of Reaction

A Study of Reaction Rates
Analyze the data using a spreadsheet.

Rates of Reaction

The Magnesium-Hydrochloric Acid System—A Study of Reaction Rates
Analyze the data using a spreadsheet.

Solutions

Limiting Reactant for a Chemical Reaction in Solution
Gather data from all students and analyze using a spreadsheet. Compute a value for each set of data. Then analyze the results collectively.

Solutions

The Freezing Point Depression of Lauric Acid
A computer-based version of this experiment is shown in the material that follows.

Stoichiometry

Stoichiometry of a Reaction and Limiting Reagent
Gather data from all students and analyze using a spreadsheet. Compute a value for each set of data. Then analyze the results collectively.

Stoichiometry

Percent Yield of a Chemical Reaction
Gather data from all students and analyze using a spreadsheet. Compute a value for each set of data. Then analyze the results collectively.

Thermochemistry

The Heat of Combustion of a Candle
Gather data from all students and analyze using a spreadsheet. Compute a value for each set of data. Then analyze the results collectively.

Thermochemistry

Heats of Reaction
Gather data from all students and analyze using a spreadsheet. Compute a value for each set of data. Then analyze the results collectively.

Transition Elements

The Amount of Copper in Copper(I) Oxide and Copper(II) Oxide
Gather data from all students and analyze using a spreadsheet. Compute a value for each set of data. Then analyze the results collectively.

Computers in Laboratory Experiments

Computers have become integral tools for data acquisition and analysis in modern chemistry laboratories. Computers may be used to obtain data, monitor data and monitor procedures during experiments. These experiments always depend upon having a detector whose output is somehow sent to the computer for analysis.

A teacher needs to attend to three things when considering the use of a computer in experiments. You must have available:

- A suitable computer. This means a computer with a "port" into which some signal can be sent from a detector. The simplest systems permit the input of a continuous signal (analog input) directly into a computer "game port."
- A detector that changes its output in response to some property of the system (*e.g.*, temperature, absorbance, intensity of emitted light) that can be connected to the computer.
- Software designed specifically for the computer and the detector.

Without all three components, computer interfacing is not possible. The three most commonly used detectors in high school chemistry experiments are:

- Photocells, that are at the heart of colorimetry experiments as well as serving as specialized detectors (*e.g.*, gas chromatography of chlorinated hydrocarbons).
- Thermistors or thermocouples that serve as temperature probes.
- pH detectors that serve in titration experiments.

Activities using these kinds of detectors have become commonplace in high school classrooms. These "computerizations" were logical extensions of typical chemistry classroom work. Early in this decade, computer applications were thought to belong in business or typing or math but not as generic tools for all areas. For this reason, there are more examples of classroom experience with hands-on computer experiments than there are with applications tools like word processors and spreadsheets, these tools having initially been placed in a different niche within the curriculum.

Many commercially available laboratory computer interfaces and software packages have become available. Some of these are excellent and economical but not inexpensive. Commercial hardware interfaces have the virtue of being ready to use off-the-shelf. At the time of this writing, the prices were at the upper level of what secondary schools typically invest in hardware. Nearly all of this hardware is accompanied by software that is extremely powerful for the tasks at hand.

In order to provide the instructions to conduct interface experiments using low cost materials, reference is made to specific systems. These references are taken from *Doing Chemistry* interfacing experiments that, in turn, were developed in collaboration with *Project SERAPHIM*.

The experiments provided here can be performed with any suitable interfaces and software. Some suggested sources of interfacing hardware and software are:

Vernier Software, 2920 S. W. 89th Street, Portland, OR 97225; (503) 297-5317

Vernier Software was founded by Dave Vernier, a teacher responding to the needs of other teachers. Vernier Software supplies software and hardware on the three common school platforms: DOS, Macintosh, and Apple. A wide range of innovative materials are offered, including some excellent manuals.

EduQuest (An IBM Educational Systems Company); (800) 426-4338

EduQuest is the source of the IBM-developed PSL laboratory materials. These interfacing materials are among the highest quality available.

Acculabs Products Group, 200 California Avenue, Suite 217, Palo Alto, CA 94306-1618; (415) 325-5898

Acculabs has developed exciting software that makes excellent use of the Macintosh interface. Hardware and software are offered.

SCI Technologies, Inc., 1716 West Main St., Suite 4, Bozeman, MT 59715; (406) 585-9088; (406) 585-8840 (FAX).

Computerized science laboratory including drop counter, Geiger counter adapter cable, pH electrode, photocell light sensor, thermistor temperature sensor and colorimeter. Both DOS-based and Macintosh-based hardware and software available.

Merlan Scientific, 247 Armstrong Avenue, Georgetown, Ontario, Canada L7G 4X6; (800) 387-2474

Merlan offers hardware and software in several platforms.

CASL Inc., 6818 86th Street E, Puyallup, WA 98371; (206) 845-7738

CASL offers complete experimental systems based upon classrooms with a few computers and electronic balances.

JCE Software, Department of Chemistry, University of Wisconsin, 1101 University Avenue, Madison, WI 53706; (608) 262-5153

JCE Software is an electronic publisher of software and has several editions including Macintosh and MS Windows.

Project SERAPHIM Catalog, Department of Chemistry, University of Wisconsin/Madison, Madison WI 53706; (608) 263-2837; (608) 262-0381 (FAX)

This catalog provides information on the construction of inexpensive computer-interfacing hardware and the accompanying software. These Laboratory Modules include instructions for making the Blocktronic (a simple colorimeter), a simple thermistor probe and a pH meter. Information on adapter boxes for the Apple II+/IIe/IIc family and the IBM-PC is also available.

Queue, Inc., 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224; (203) 335-0906; (203) 336-2481 (FAX)

The Queue Science Catalog provides information on hardware and software for interfacing Apple II+/IIe/II-GS and IBM-PC computers.

LABORATORY ACTIVITY: STUDENT VERSION

Laboratory Activity 1: Cooling Curves

(Modified from Experiment F23 of *Doing Chemistry*)

Introduction

When a substance melts or freezes, the temperature of the solid/liquid mixture remains constant even as heat is added/removed.

Purpose

To study the energy changes associated with cooling and heating a substance through a phase change.

Safety

Burns are possible. Handle heating equipment cautiously. Check glassware by holding finger nearby to sense heat before touching. Wear goggles and apron.

Procedure

1. Install a pair of thermistors in the appropriate interface equipment.
2. Boot the computer using appropriate software. Choose a sampling time interval of about 7 sec. Select graphic display. Adjust the maximum and minimum temperatures.
3. Use a tripod, gauze, beaker, water, and burner, or use a hot plate to prepare a water bath at 60-70 °C. Place 10 g of lauric acid in a clean, dry 25- x 150-mm test-tube. Heat in a hot water bath to a temperature of 50-60 °C until all of the solid crystals have melted.
4. Prepare a calorimeter. Insert a 600-mL beaker into a 1-L beaker containing crumpled newspaper or paper towels for insulation. Place 400 mL of tap water into the inner beaker. Prepare a cardboard cover. Insert one thermistor into this cover. The cover should have a hole large enough to accommodate the test-tube. Place the cover over the beaker. Prepare a thermistor probe by using a small rubber band to attach the second thermistor to a glass stirring rod. Remove the melted lauric acid from the hot water bath. Insert the thermistor probe. Place through the cardboard cover into the calorimeter. Stir gently.
5. When the setup is complete, select the option to begin data collection. Stir continuously during data collection. A set of axes scaled according to your specifications will appear. Data points will be plotted for the selected time interval, every 7 sec in this case. Gather data until the two probes indicate the same or nearly the same temperature.
6. Save the data to a computer file.
7. A heating curve also can be studied. Remove the tube from the calorimeter. Discard the water. Heat 300 mL of water to a temperature of 70 °C. Pour this hot water into the calorimeter. Replace the tube of solidified lauric acid, and start recording a new set of temperatures.
8. Save the data when requested.
9. Wash your hands thoroughly before leaving the laboratory.

Data Analysis and Concept Development

Provide an interpretation of the screen graphs. How is it possible that one graph has bumps while the other one changes smoothly? Account for the differences in total temperature change between the two graphs.

LABORATORY ACTIVITY: TEACHER NOTES

Laboratory Activity 1: Cooling Curves

Major Chemical Concept

The temperature does not change even though heat is added or removed when the phase changes for a substance.

Level

This is part of the mainstream high school chemistry class, but a nonquantitative version can be included in lower-level chemistry classes.

Expected Student Background

The concepts of heat and temperature should precede this experiment.

Time

Teacher preparation: 15 min
Procedure: 40 min

Safety

Be certain to use substances that do not pose a special toxicity problem. Lauric acid is a good choice. Also, since it is an acid, it can be washed up from glassware using a mild base such as sodium bicarbonate. Emphasize use of goggles and apron when performing laboratory activities.

Materials (For 24 students working in pairs)

Nonconsumables

Suitable computers
Suitable interfaces with two calibrated temperature probes for each computer
Suitable computer software
12 Beakers, 400-mL
12 Beakers, 600-mL
12 Burners, igniter
12 Test-tubes, 25- x 150-mm
12 Thermometers, –10 to 110 °C
12 Glass stirring rods, 20-cm length

Consumables

180 g Lauric acid (or other suitable substance)
Paper towels

Advance Preparation

All hardware should be prepared in advance.

Pre-Laboratory Discussion

Get the students to commit to a position about changing the temperature when adding or removing heat. Most will insist that the temperature must change when heat is added or removed. Get students to commit to their idea, correct or not, before the experiment starts.

Teacher-Student Interaction

Most of your time in the laboratory will be spent helping students with both the hardware and software. The more computer experiments you have in place, the less serious will be this kind of work. While you help, it is good to point out what the computer is actually doing during this experiment. In a temperature experiment, the computer is recording pairs of data points, usually time and either temperature or a number that is related to temperature such as a voltage.

Anticipated Student Results

Typical student results are displayed in the figure.

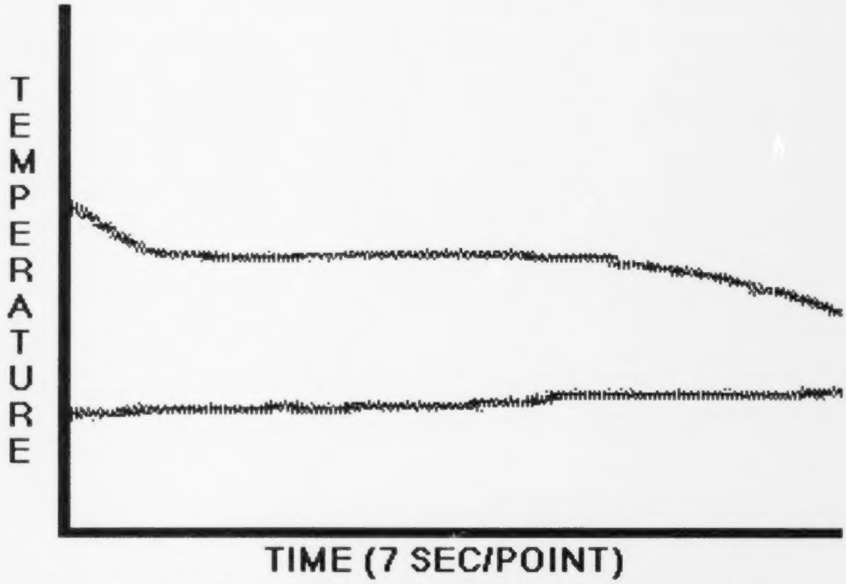


Figure 11. Typical cooling curve shown on computer screen.

Answers to Data Analysis and Concept Development

When a substance undergoes a phase change, all of the added (removed) energy goes into changes in potential energy and none goes into changing the average kinetic energy (which, if it did change, would be reflected in a temperature change.)

Post-Laboratory Activities

Discussion of the results is very important. Heat is gained at a smooth continuous rate by the calorimeter bath fluid. This is indicated by the smooth upward drift of the temperature of the bath fluid, and this is consistent with most students' intuition. Heat is also lost at a smooth, continuous rate by the freezing liquid. However, the temperature curve shows regions of rapid temperature change (steep changes) and ones of no temperature change. This behavior seems counterintuitive to most students. Use the post-laboratory to compare results and interpretations of the data with predictions made by students.

Ask students to explain the effect that performing this experiment with two temperature probes instead of just one temperature probe had on their thinking. What would they have predicted for the outside temperature change if all they had to go by were the inside temperature change? What would they have guessed for the inside temperature change if all they had were the outside temperature change?

Assessing Laboratory Learning

1. Sketch the curves you would expect to measure if the solid lauric acid at room temperature were placed in a calorimeter bath at a temperature of 90 °C.
2. Discuss the statement, "The cooling curve for the freezing of a liquid is a characteristic "fingerprint" that is extremely useful when identifying that substance."

LABORATORY ACTIVITY: STUDENT VERSION

Laboratory Activity 2: Equilibrium Constant

(Modified from Experiment F45 of *Doing Chemistry*)

Introduction

Some chemical reactions do not proceed completely; instead, after the reactants have been reacting for some time, they come to a "chemical equilibrium mixture" containing both unreacted chemicals and reaction products.

Purpose

To study an equilibrium reaction and to discover quantitative relationships between the amount of reactant and product at chemical equilibrium.

Safety

Thiocyanate is poisonous if ingested. Concentrated nitric acid will stain or eat away hands and clothes and is toxic. Do not ingest chemicals. Wash hands before leaving class. Wear laboratory aprons and goggles.

Procedure

1. Follow the procedure for the particular computer, interface, and software available.
2. Connect the interface to the computer. Boot the computer using appropriate software.
3. Use a calibrated plastic pipet to place 4 mL of 0.3 M nitric acid into each of six labeled 13- x 100-mm test-tubes (Tube 1 to Tube 6). Use a calibrated plastic pipet to place 8 mL of 0.2 M iron(III) nitrate into another test-tube (Tube 0).
4. Use the pipet to withdraw 4 mL of iron(III) nitrate solution from Tube 0 and place it in the first labeled tube of acid. Use the pipet to mix carefully the iron solution and acid. Then withdraw 4 mL from the first tube and add to the second tube. Repeat for all tubes. Discard the 4 mL removed from the sixth tube.
5. Use a fresh plastic pipet to add 4 mL of 0.001 M potassium thiocyanate to each of the seven tubes. Select a small square of waxed paper. Use the thumb to cover a tube with this paper. Invert to mix tube contents. Repeat for all tubes.
6. Place distilled water in a clean test-tube that fits into the probe. Insert the tube into the colorimeter probe, and cover with a light shield. Use this water to adjust the probe.
7. Remove the light shield. Discard the water from the probe test-tube. Select the most dilute solution. Rinse the probe test-tube with a small portion of this solution. Discard the rinse. Pour the remaining solution into the tube. Replace the tube in the probe. Cover the tube with a light shield.
8. Follow the instructions (usually available on the computer screen) to obtain a relative absorbance reading, and record that reading from the computer screen.
9. Repeat the measurement for each of the solutions proceeding from the least to the most colored.
10. Wash your hands thoroughly before leaving the laboratory.

Data Analysis and Concept Development

1. Assume that, for the highest iron concentration, *all* of the SCN^- is converted into complex ion. From the absorbance measured under these conditions, determine a constant related to absorbance and the amount of complex ion formed.
2. Use the experimental data to determine the concentration of complex present, the amount of unreacted SCN^-, the amount of unreacted Fe^{3+} present for each solution measured. One way to determine whether a quantitative relationship exists between the concentrations of reactants and products is to develop functions involving the concentrations and test them numerically.
3. Calculate the value of the following expressions for each data set:

I. $$\frac{[FeSCN^{2+}]_{eq}}{[Fe^{3+}]_{eq} + [SCN^-]_{eq}}$$

II. $$\frac{[FeSCN^{2+}]_{eq}}{[Fe^{3+}]_{eq}{}^{2}}$$

III. $$\frac{[FeSCN^{2+}]_{eq}}{[Fe^{3+}]_{eq} \times [SCN^-]_{eq}}$$

IV. $[Fe^{3+}]_{eq} \times [SCN^-]_{eq} \times [FeSCN^{2+}]_{eq}$

Analyze the resulting data to decide which, if any, of these relationships leads to a constant numerical value.

Laboratory Activity: Teacher Notes

Laboratory Activity 2: Equilibrium Constant

Major Chemical Concept

Chemical equilibria are important and may be studied quantitatively.

Level

Equilibrium is covered in about half of the high school chemistry courses at the end of the year. Equilibrium is major topic in advanced placement chemistry.

Expected Student Background

Students should practice with spreadsheets before attempting this experiment. Prior topics include stoichiometry and (usually) reaction kinetics.

Time

Teacher preparation: 1 hr
Procedure: 50 min
Data analysis: 50 min

Safety

Thiocyanate is poisonous if ingested. Concentrated nitric acid will stain or eat away hands and clothes and is toxic. Warn students not to ingest chemicals. Have students wash their hands before leaving class. Warn students to wear older clothing or laboratory aprons.

Materials (For 24 students working in pairs)

Nonconsumables

- 180 Beakers, 50-mL
- 60 Test-tubes to fit a colorimeter
- 420 Test-tubes, 13- x 100-mm
- 120 Calibrated 4-mL (or 3-mL) plastic dropping pipet (other dilution apparatus will work equally well; this is a particularly effective, low-cost choice)
- 12 Test-tube racks
- Suitable computer
- Suitable computer interface
- Suitable software
- (Overhead projector, 100-mL beaker, two 10-mL graduated cylinders with detachable plastic bases, Pasteur pipet)

Consumables

- 3000 mL 0.2 M Iron(III) nitrate [243.6 g $Fe(NO_3)_3 \cdot 9H_2O$, plus 60 mL concentrated HNO_3 diluted with enough distilled water to make 3000 mL of solution]
- 2400 mL 0.0010 M Potassium thiocyanate (see *Advance Preparation*)
- 2400 mL 0.3 M Nitric acid (240 mL concentrated nitric acid diluted to 2400 mL with distilled water)
- 420 Waxed paper, 3- x 3-cm squares

Advance Preparation

Check this experiment in advance; many things can go wrong.

It is less expensive to add this experiment using suitable interfaces to a computer (if one is available) than to purchase a spectrophotometer.

Prepare a 0.001 M solution of potassium thiocyanate in two steps. Weigh 9.71 g KSCN, and use distilled water to make 1 L of solution (0.10 M). Take a 10 mL aliquot of this solution, and dilute to 1 L with distilled water (0.0010 M). Do this step three times to prepare 3000 mL.

Prepare the iron(III) nitrate solution using the dilute nitric acid as solvent instead of water. The iron(III) nitrate solution should be nearly colorless.

Disposable calibrated plastic dropping pipets of 3-mL or 4-mL volume are available from some supply houses and work very well for this experiment. They are very low in cost and are reusable many times. Have students practice mixing solutions with the pipets before they begin the experiment.

With just a bit of practice, fitting data using a spreadsheet can be accomplished. In the case shown here, data were fitted by using one cell in which the concentration factor was set arbitrarily. The first concentration factor used was 0.0005/0.804, the molar concentration of thiocyanate divided by the absorbance. This factor was used to calculate all of the concentrations of $FeSCN^{2+}$, which in turn was used together with the initial concentrations to determine the equilibrium concentrations. To find a best fit, the values for the equilibrium constant were averaged and added the absolute difference between each individual value and the average. The sum of the absolute differences was minimized by changing the concentration of $FeSCN^{2+}$ found for the first case. This number can be 0.0005 M as a maximum, but will be less if the reaction is not 100% complete. The minimum value for the sum of the differences was found when the concentration of $FeSCN^{2+}$ in the first experiment was set to 0.000481 instead of 0.0005 M.

Pre-Laboratory Discussion

Discuss the notion of a reaction equilibrium, and the interest in determining quantitative relationships between concentrations of reactants and products.

Teacher-Student Interaction

Much of your work as you walk around the laboratory will be in helping students to get their hardware and software functioning. This problem decreases as they perform more and more experiments using computers. While students are using spreadsheets (such as AppleWorks) work with them to be certain that they understand how the spreadsheet functions and how the computations are being made.

Anticipated Student Results

Initial Concentrations

$[SCN^-]$	$[Fe^{3+}]$	Absorbance
0.000500	0.100	0.804
0.000500	0.0500	0.770
0.000500	0.0250	0.683
0.000500	0.0125	0.623
0.000500	0.00625	0.511
0.000500	0.00313	0.386
0.000500	0.00156	0.258

Figure 12. Typical absorbance data for iron(III) and thiocyanate solutions.

Answers to Data Analysis and Concept Development

The data above are analyzed assuming that all of the SCN^- reacts to form $FeSCN^{2+}$ in the 0.10 M Fe^{3+} mixture. The absorbance is then used to calculate the $[FeSCN^{2+}]_{eq}$ for the other reaction mixtures.

$$\text{Constant} = 0.0005/0.804 = 6.22 \times 10^{-4} \text{ M/absorbance reading}$$

Using this constant and the balanced chemical equation,

$$Fe^{3+} + SCN^- \rightleftharpoons FeSCN^{2+}$$

calculate the equilibrium concentrations for all of the species in solution at equilibrium:

$$FeSCN^{2+}]_{eq} = 6.22 \times 10^{-4} \text{ M} \times \text{Absorbance}$$

$$[Fe^{3+}]_{eq} = [Fe^{3+}]_{init} - [FeSCN^{2+}]_{eq}$$

$$[SCN^-]_{eq} = [SCN^-]_{init} - [FeSCN^{2+}]_{eq}$$

The results are shown in Figure 13.

$[SCN^-]_{init}$	$[Fe^{3+}]_{init}$	Absorb	$[FeSCN^{2+}]_{eq}$	$[Fe^{3+}]_{eq}$	$[SCN^-]_{eq}$
0.000500	0.100	0.804	5.00E-04		
0.000500	0.0500	0.770	4.79E-04	4.95E-02	2.11E-05
0.000500	0.0250	0.683	4.25E-04	2.46E-02	7.52E-05
0.000500	0.0125	0.623	3.87E-04	1.21E-02	1.13E-04
0.000500	0.00625	0.511	3.18E-04	5.93E-03	1.82E-04
0.000500	0.00313	0.386	2.40E-04	2.89E-03	2.60E-04
0.000500	0.00156	0.258	1.60E-04	1.40E-03	3.40E-04

Figure 13. Equilibrium concentrations (E = Exponent).

Use a spreadsheet to examine different extents of reaction. When fitted in this fashion, the data for all seven tubes may be included. When the equilibrium concentration of $FeSCN^{2+}$ in the 0.10 M Fe^{3+} data set is assumed to be 0.000481, the following results (Figure 14) are obtained.

$[SCN^-]_{init}$	$[Fe^{3+}]_{init}$	Absorb.	$[FeSCN^{2+}]_{eq}$	$[Fe^{3+}]_{eq}$	$[SCN^-]_{eq}$	Ratio
0.000500	0.100	0.804	4.81E-04	9.95E-02	1.90E-05	254
0.000500	0.0500	0.770	4.61E-04	4.95E-02	3.93E-05	236
0.000500	0.0250	0.683	4.09E-04	2.46E-02	9.14E-05	182
0.000500	0.0125	0.623	3.73E-04	1.21E-02	1.27E-04	241
0.000500	0.00625	0.511	3.06E-04	5.94E-03	1.94E-04	265
0.000500	0.00313	0.386	2.31E-04	2.90E-03	2.69E-04	296
0.000500	0.00156	0.258	1.54E-04	1.41E-03	3.46E-04	318

Figure 14. Spreadsheet of data for all seven tubes.

When one "fits" the data, the ratio has an average of 256 and a 1.75-fold range. It is the same as saying that only 96.2% instead of 100% of the thiocyanate is tied up in the complex.

$$K_c = \frac{[FeSCN^{2+}]}{[Fe^{3+}]\,[SCN^-]}$$

For either of these cases (raw or "fitted" data) the other expressions that were tried had a much larger variation.

2. A table of values for the expressions I, II, III, and IV, using data for the first through seventh experiments based upon best fit for ratio III is given in Figure 15.

I	II	III	IV
4.8E-03	0.0500	254	9.10E-10
9.3E-03	0.190	236	8.98E-10
1.7E-02	0.680	182	9.18E-10
3.0E-02	2.53	241	5.75E-10
5.0E-02	8.65	265	3.53E-10
7.3E-02	27.5	296	1.80E-10
8.8E-02	78.1	318	7.50E-11

Figure 15. Values for expressions I-IV using data from all experiments.

Post-Laboratory Activities

Discussion of the results and verification that students have done and understood their calculations is important.

Assessing Laboratory Learning

1. Predict the effect of a 1% error in the reading of the absorbance of the first tube. Say the true reading should have been 0.812. Show the impact of that error on the equilibrium constant calculation.

   Answer:

$[SCN^-]_{init}$	$[Fe^{3+}]_{init}$	Absorb.	$[FeSCN^{2+}]_{eq}$	$[Fe^{3+}]_{eq}$	$[SCN^-]_{eq}$	Ratio
0.000500	0.100	0.804	0.000495	0.0995	4.93E-06	1010
0.000500	0.0500	0.770	0.000474	0.0495	2.59E-05	370
0.000500	0.0250	0.683	0.000421	0.0246	7.95E-05	215
0.000500	0.0125	0.623	0.000384	0.0121	0.000116	272
0.000500	0.00625	0.511	0.000315	0.00594	0.000185	286
0.000500	0.00313	0.386	0.000238	0.00289	0.000262	314
0.000500	0.00156	0.258	0.000159	0.00140	0.000342	332

Figure 16. Calculations for impact of error on equilibrium constant calculation.

LABORATORY ACTIVITY: STUDENT VERSION

Laboratory Activity 3: Candles

Introduction

The time required for a candle to extinguish in a restricted volume is studied quantitatively.

Purpose

Even simple experiments provide students with opportunities to use a wide array of computer productivity tools.

Safety

Burns are possible. Handle heating equipment cautiously. Check glassware by holding finger nearby to sense heat before touching. Have moist towels and a fire extinguisher available. Goggles and aprons are required.

Procedure

1. Place a small laboratory candle on a flat surface such as a piece of cardboard. Light the candle. Use a watch with a seconds sweep hand or a seconds digital readout. Note and record the time while placing an inverted beaker over the burning candle. Note and record the time that the flame extinguishes. Determine the duration for which the candle burns under the beaker. Repeat the experiment using different sized beakers and different sizes and lengths of candles.
2. Predict the time for burning of a beaker for which that time has not been measured. Measure the time for that beaker. Compare the predicted and observed times.

Data Analysis and Concept Development

Prepare a report for this experiment using a word processing applications program. Determine the times of burning using a spreadsheet in which the start and stop times are entered into the cells of the sheet, and the flame duration time is computed from them. Use a graphing program to prepare a plot of the flame duration versus beaker volume for each candle size and shape tested. Use a paint program to prepare an illustration of the experiment for inclusion within your report.

Develop an hypothesis to explain the quantitative results.

LABORATORY ACTIVITY: TEACHER NOTES

Laboratory Activity 1: Candles

Major Chemical Concept

Computer applications are useful in preparing reports for all experiments.

Level

Introductory chemistry, all levels.

Expected Student Background

No special skills required. Students should have had experience with all of these applications programs in other classes, else concurrent instruction in program use should be provided.

Time

Teacher preparation: 5 min
Procedure: 50 min

Safety

Burns and fires are possible with this experiment. Have moist towels and a fire extinguisher available for use. Remind students of the fire and burn hazards at the beginning of the experiment.

Materials (For 24 students working in pairs)

Nonconsumables

- Beakers in various sizes (150-mL, 250-mL, 400-mL, 600-mL, 1000-mL); it would be good to have beakers in families of shapes such as standard and tall-form shapes.
- 12 Timing devices (students' watches; wall clocks)
- Computers
- Applications programs (word processing; spreadsheet with graphing or with companion graphing; paint/draw)

Consumables

- 24-36 Laboratory candles
- Matches
- 12 Cardboard squares

Advance Preparation

Mount the candles on the cardboard squares or other surfaces.

Pre-Laboratory Discussion

There is a traditional experiment developed in the CHEM Study project that involves making observations about a burning candle. This experiment is included early during the semester, often on the first or second day of class. The activity described here is a way to extend this first-day-of-class activity to include systematic computer applications for the activity.

Teacher-Student Interaction

While the students are gathering data, encourage them to explore a different range of possibilities, and to repeat experiments to confirm results.

Anticipated Student Results

No simple rules are available to predict the length of burning time. The larger the beaker, the longer the burning time. The shorter the candle, the longer the burning time.

The data shown in Figures 17-19 were obtained using typical laboratory beakers.

Volume (mL)	Start	Stop	Difference	Duration (sec)
100	11:00:17	11:00:20	0:00:03	3
150	11:02:06	11:02:12	0:00:06	6
150	11:03:30	11:03:37	0:00:07	7
250	11:06:01	11:06:10	0:00:09	9
250	11:08:01	11:08:11	0:00:10	10
1000	11:11:30	11:11:56	0:00:26	26
1000	11:14:00	11:14:30	0:00:30	30
400	11:17:30	11:17:45	0:00:15	15
400	11:20:10	11:20:24	0:00:14	14

Figure 17. Sample data.

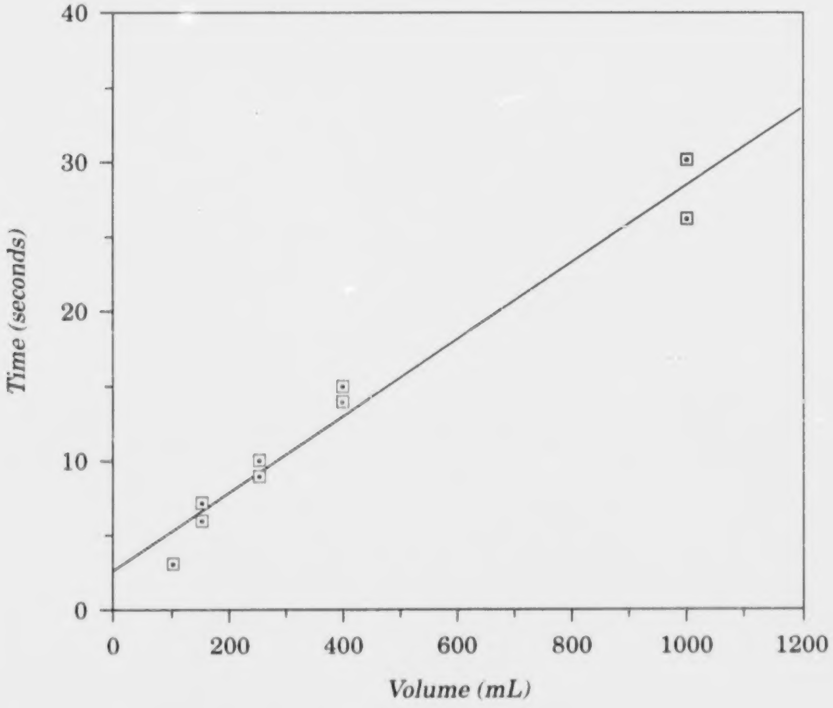


Figure 18. Experimental data fitted with linear equation.

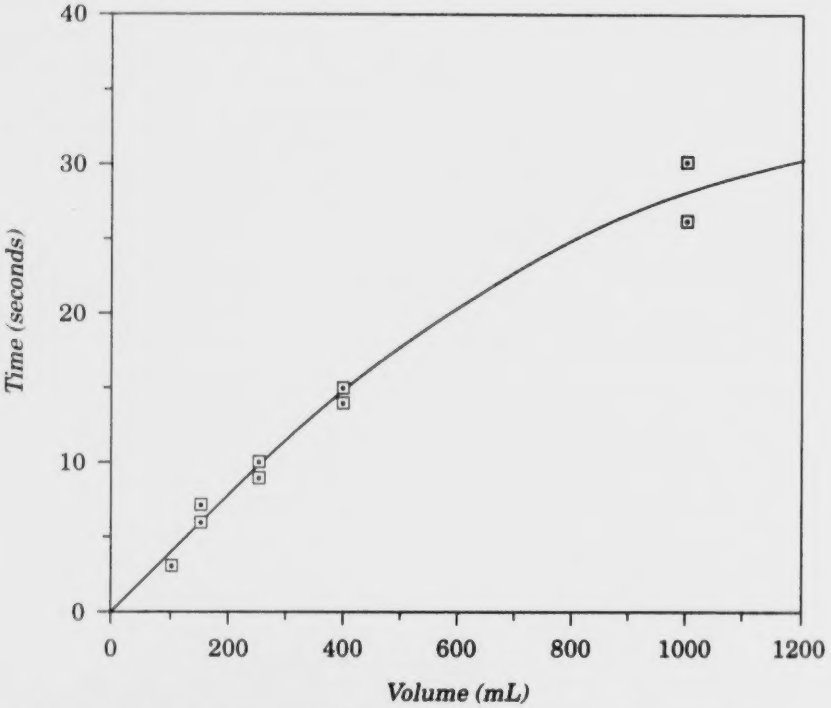


Figure 19. Experimental data fitted with parabolic equation.

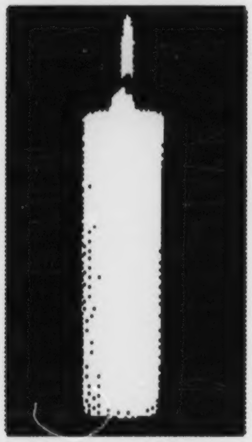


Figure 20. Use of paint program to illustrate image of burning candle.

Answers to Data Analysis and Concept Development

This is more of a real-world experiment than a "pretty" experiment with simple explanations and clear cut answers. It is a particularly good experiment to start students with an empirical notion of how difficult it is to solve some seemingly simple problems.

Post-Laboratory Activities

This experiment often works well when students are assigned to teams. Each team is charged with making comprehensive observations, recording them, and preparing a report. The quality of the reports may be enhanced in several ways. Assigning numbers to the groups, choosing group numbers randomly from numbered slips drawn from (but returned to) a fish bowl, and then having the group report and justify an observation and interpretation enhances student performance.

Have each group prepare a separate report using all of the computer tools indicated. Then share the reports and ask each group to use their electronic files to develop a "combined report." Finally, ask the class to come up with one single class report that pulls together their best consensus as a class unit.

Assessing Laboratory Learning

There are several ways in which the results can be assessed. The quality of the reports can be assessed. There can be a best group idea, best group initial report, best graphical analysis, best consensus report, best idea for a continuing experiment, and other "best" awards based upon the reports.

References

This module was developed by Robert Curtright, James McGahan, and David Brooks.

Holmquist, D. D., & Volz, D. L. (1991). *Chemistry with computers.* Vernier Software, 2920 S.W. 89th Avenue, Portland, OR 97225. (503) 297-5317.

Thirty-one chemistry student experiments and teacher demonstrations using Vernier Software products with Apple and IBM computers for collecting, displaying, printing, graphing and analyzing data.

Moore, J. W. (Ed.). Bits and pieces. *Journal of Chemical Education.*

Each month this column on computers publishes information on software developed by the authors.

Project SERAPHIM. *Catalog.* Department of Chemistry, University of Wisconsin, Madison, WI 53706. (608) 263-2837; (608) 262-0381 (FAX).

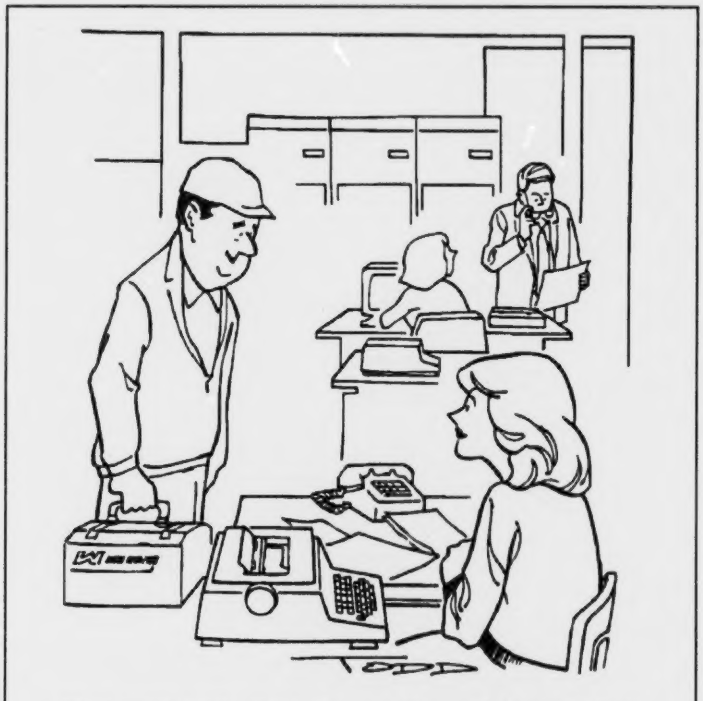
Software an materials for computers in chemistry instruction. It is free to members of Project SERAPHIM; the cost is $10 to nonmembers. It lists materials available from the SERAPHIM Clearinghouse: Written Materials for teaching, programming and laboratories; Software (about 650 different programs) for Apple II, Apple Macintosh, Apple II ProDOS and IBM-PC/PC compatibles.The appropriate programs are referenced in the individual *SourceBook* content modules.

Falcon Software, Inc. *Catalog.* Box 200, Wentworth, NH 03282. (603) 764-5788.

This catalog lists software and multimedia (CD-ROM packages, videodisks, *etc.*). The appropriate programs are referenced in the individual *SourceBook* content modules.

Appendix

- **Humor**

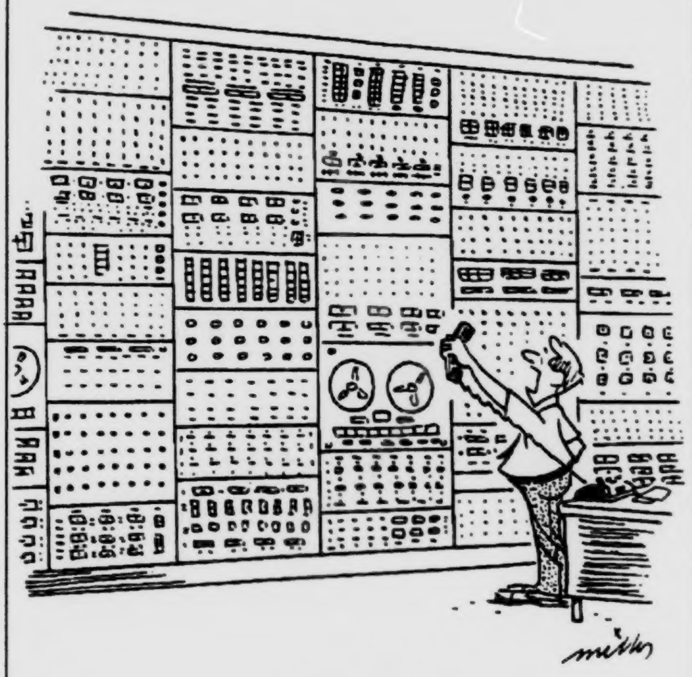


"Gud morning. Wood you show me the werd-procassor thats bin ackting op?"

CHEMTECH, October 1984, p. 579. ©American Chemical Society. Reprinted with permission.

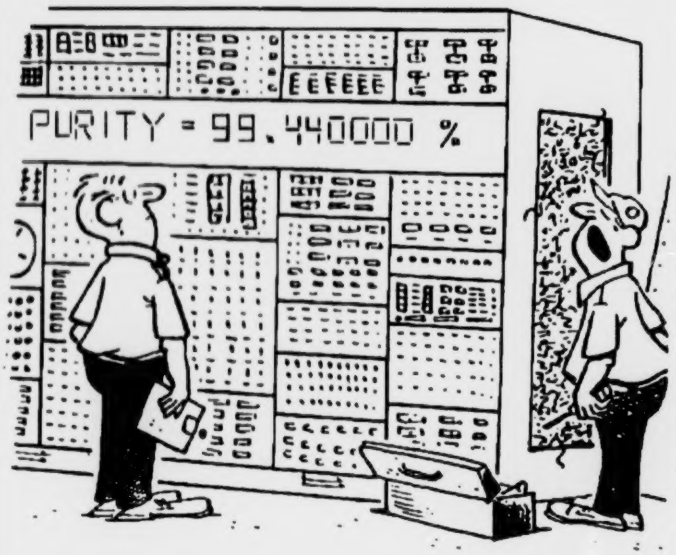
"The poor dear has been that way ever since he stepped on a floppy disk he hadn't backed up."

CHEMTECH, September 1984, p. 517. © American Chemical Society. Reprinted with permission.



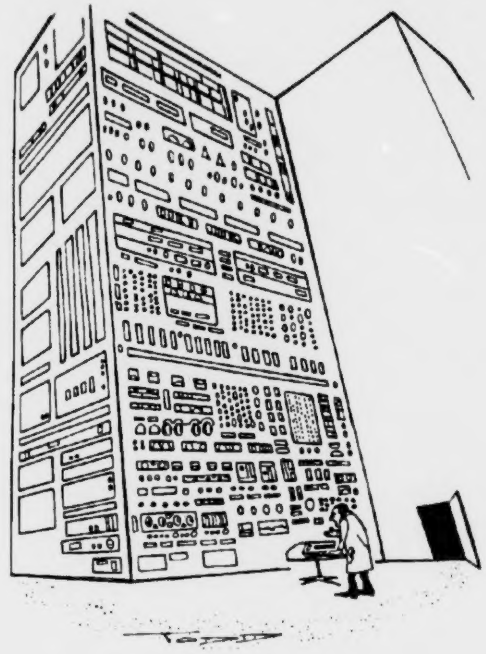
"It's for you!"

CHEMTECH, January 1980, p. 57. © American Chemical Society. Reprinted with permission.


PURITY = 99.440000 %

"Looks like the best she can do, boss."

CHEMTECH, January 1980, p. 22. © American Chemical Society. Reprinted with permission.



"It's me again . . . I hate to be a pest, but would you mind answering just one more little question?"

CHEMTECH, August 1980, p. 462. © American Chemical Society. Reprinted with permission.

CHEMISTRY IN YOUR CAREER

A SourceBook Module

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TRADITIONAL AND NON-TRADITIONAL CAREERS

A knowledge of chemistry is a valuable asset for a variety of careers. These careers include archaeology, agriculture, chemical engineering, science writing, forensics, museums, chemical sales, patent law, information science, banking, industrial hygiene, toxicology, medicine, dentistry, epidemiology, pharmacology, government, insurance, and computers, among others.

An education in chemistry is valuable in entering these areas for many reasons. Chemistry can provide the factual background needed to explain science news for the general public (science writing) or help understand the use of science in novels (science fiction writing). It can help contribute to efficient functioning of the criminal justice system by providing objective scientific evidence (forensics) or help discover the technologies of past ages, how to restore paintings and sculpture, how to date art and artifacts, and how to determine the optimum environmental conditions for displaying art and artifacts (museums).

A background in chemistry can lead to opportunities in the chemical business ranging from making chemicals to selling and/or buying them. Additionally, the chemical business graduate studies the market for various products, arranges financing for chemical endeavors, builds plants, and tests products for quality. Chemists commonly move into positions to manage the efforts of people doing work in the chemical industry.

A chemistry background is useful to the field of patent law where a client would benefit from an individual's or an attorney's knowledge about science and inventions. With a chemistry background, an industrial hygienist can monitor company employees to minimize exposure to harmful chemicals or environmental stresses; a toxicologist can ascertain the potentially adverse effects of chemicals in biological systems.

Careers in health-related areas such as medicine and dentistry have traditionally required a strong background in chemistry since chemistry is such an intrinsic part of these disciplines. None of the careers discussed is devoted specifically to chemistry, but a knowledge of chemistry is important in its own way for each career. The careers described above are often called *nontraditional fields or careers* because prior chemistry training permits graduates to pursue productive and successful careers in areas outside of the mainstream of chemistry.

There are also many careers that are directly related to chemistry and require persons designated as chemists *(traditional careers)*. Chemists are generally categorized according to the five traditional branches of chemistry:

- Analytical chemists determine the composition of substances, measure trace components, and separate complicated mixtures into their components. More importantly, they develop new methods of chemical analysis.
- Biochemists study the reactions and mechanisms that occur in plant and animal cells, and methods for preparing biologically useful compounds. Biotechnologists must be well versed in both analytical chemistry and biochemistry.

- Inorganic chemists investigate a variety of compounds formed from all the known elements.

- Organic chemists study the preparation, reactions, mechanisms, structures, and properties of natural and synthetic carbon compounds.

- Physical chemists seek quantitative explanations in precise mathematical terms which describe and predict the physical properties and reactions of atoms and molecules.

Chemists work in research and development (discovery and production of substances in commercial quantities in usable form at acceptable cost) or teach in public or private schools. Some chemists work in government agencies, or pursue careers in the business and management components of industrial chemical corporations.

Regardless of the professional goal, education for a career described above should begin in high school. The high school curriculum should provide basic skills in English and speech, mathematics, science (at least or a minimum of one course each in biology, chemistry, and physics), social science, and computer literacy (the use of the computer as a tool).

TRAINING

There is a wide variety of chemistry programs available at the post secondary level. Many colleges and universities offer a course for non-science majors which provides an introduction to understanding chemical involvement in our daily lives such as maintaining clean air and water, hazardous waste disposal, and awareness of chemicals in our daily lives in medicinals, cosmetics and fuels. This type of course is designed for students who have not decided on a field of study and are interested in exploring science-related options.

If one chooses to major in chemistry, the college curricula are typically divided into four groups:

- General Studies - courses from the humanities, social sciences, fine arts, and communication

- The Major - a minimum of one chemistry course each semester including general chemistry, organic chemistry, physical chemistry, and analytical chemistry. Courses in inorganic, nuclear, biochemistry and polymer chemistry may also be required. A research experience and/or other advanced courses are often required or may be elected.

- The Minor - six to eight courses in another area, *e.g.*, physics, mathematics, biology or another area of study outside of science. Similarly, a chemistry minor is valuable support for a number of other majors (biology, nursing, allied health professions, *etc.*).

- Electives - courses selected from any area of study.

- Chemistry majors are also required to take mathematics at least through calculus, and one year of calculus-based physics.

A major or minor in chemistry opens the door to the wide variety of professions utilizing chemistry. Some popular professions include:

• Biotechnology - This subdiscipline of chemistry utilizes biological processes to learn about the science of life and to find ways to manufacture new materials. Biotechnologists attempt improvement in diagnosis and treatment of hereditary diseases, tackle environmental problems such as pollution, investigate ways to clone genes or cells with desirable characteristics, and study ways to improve crop yield, and the nutritional value of food. *Job titles include molecular biologist, genetic engineer, immunologist, biochemist, and geneticist.*

• Chemical Education - Teachers, like parents, influence a young person's development. They are educators, role-models, and mentors. Teaching is stimulating and demanding, requires a sound knowledge of chemistry and daily interaction with students. Career opportunities for chemists include teaching in elementary and high schools, two- and four-year colleges and universities. *Job titles include professor, associate professor, assistant professor, instructor, chemistry teacher, science teacher, science supervisor, and district science coordinator.*

• Chemical Engineering - Chemical engineers apply scientific and mathematical principles to solve practical problems. They translate processes developed on the small scale in the laboratory into practical methods for large scale production. Plant design, evaluation of plant operations and product quality, laboratory research, economic design, customer service, plant production and environmental studies are examples of work performed by chemical engineers. *Job descriptions include production supervisor, plant or project engineer, design contractor, heat-transfer engineer, development engineer, and environmental engineer.*

• Materials Science - Materials Science is an applied science and deals with the relationship between structure and properties of materials. This information is then used to make new materials with special properties, for example, superconductors, strong and lightweight, flexible and resilient graphite materials for car bodies and tennis racquets, and fuel cells for spacecraft use. *Job descriptions are materials specialist, ceramist, metallurgist, and polymer chemist.*

• Medicine and Dentistry - Practitioners in health fields are concerned with why and how certain drugs are effective against certain diseases or what a particular diagnostic test means. The experimental methods learned while studying chemistry and biology help to develop the ability to solve problems logically and objectively. *Job descriptions include practicing physician, or dentist, and research physician or dentist.* In addition to the traditional allopathic medical fields (oncology, cardiology, orthopedics, *etc.*), the term *practicing physician* includes optometry, veterinary medicine, podiatry, osteopathy, chiropractic, *etc.*.

Graduate education in chemistry is essential for many of the fields outlined above. A master's degree is the beginning of specialization in an area of chemistry. At this level, the student spends time concentrating in one area or specialization in preparation for a position in that particular area. Required curricula are course work and an original research project. A doctoral degree in chemistry is the highest attainable professional degree. It provides additional concentration in one or more areas as well as independent, fundamental research. The doctoral degree is required in nearly forty percent of the professional positions in chemistry, from research chemist in a large corporation to university/college professor.

CHOOSING A CAREER

Thus, there is a wide variety of careers available utilizing chemistry. These career choices depend essentially on three factors: (1) the kind of work a person likes to do - research and development, production, testing, teaching, or management, (2) the area the person would like to work in - private, academic, government, and (3) the specific areas of chemical knowledge being investigated.

It is often not easy to decide which career to follow. A few recommendations that will make the decision easier are:

- Determine your interests and attitudes
- Seek help from guidance counselors or college program advisors
- Ask questions about specific careers
- Explore a broad variety of topics beginning with the general (*TIME*, *Newsweek*, newspapers, *etc.*) and work toward the more specific (journals, periodicals, advanced texts, *etc.*)

It is important not to set goals too low. Students often change career goals several times. Unfortunately, many career paths are effectively closed if students set their academic goals too low at an early stage of their training. It is not uncommon to decide on a career in chemistry after obtaining a four-year college degree in another field. Therefore, it is important to keep an open mind concerning the selection of courses you take in high school and as a college undergraduate. Remember, over a lifetime interests can change, and a person's interests at age 20 are not likely to be the same as they will be at age 35 or 50.

The following list contains the ideas to keep in mind in preparing for a career while you are a college undergraduate:

- Develop a background as broad as possible for as long as possible to take advantage of as many opportunities as possible.
- Get involved in related opportunities as they arise.
- Seek a variety of experiences.
- Start early to prepare a resumé and regularly update it.
- Develop healthy work habits early on – a job at a fast food restaurant may not be the ultimate goal, but habits developed early can easily affect an entire career.
- Be prepared from the beginning. Interviews and first impressions are often long-lasting.

STUDENT AFFILIATES PROGRAM

The Student Affiliates Program introduces undergraduate students in the chemical sciences to the benefits of professional association and to the principles and purposes of the American Chemical Society (ACS). Benefits to student affiliates include a journals option package; *in Chemistry*, a newsletter published five times per year; low rates on subscriptions to ACS journals; and use of the Employment Clearing House. Benefits to chapters include a film/video library and portfolios of chapter materials. Affiliates can also participate in undergraduate research presentations at ACS spring national meetings. A booklet entitled "Student Affiliates/Local Section Activities" provides ideas and suggestions for increased interaction between student affiliates chapters and local sections. See address under resource list.

RESOURCE LIST

The following ACS Materials, available by writing to the following, are useful career information resources:

Career Publications
American Chemical Society
1155 16th Street, N.W.
Washington, D.C. 20036
(202) 872-4600

1. *Careers in Chemistry: Questions and Answers;* 1988, 32 pp.
2. *Careers in Chemistry: Chemistry in the Community*; 1989, 10 pp.
   Highlights eight careers in which chemistry plays a role.
3. *A Career as a Chemical Technician;* 1988, 6 pp.
4. *Careers in the Life Sciences*; 1984, 19 pp.
5. *Employment Outlook*; published annually, 18 pp.
   A reprint from *Chemical and Engineering News* that gives up-to-date perspectives on the employment outlook for chemists and chemical engineers. Prospects for women and minorities are included as well as the latest salary data.
6. *Futures Through Chemistry*; 1985, 12 pp.
   This booklet is intended to help college students planning a chemistry career refine their general interest in chemistry into specific educational and career objectives—to help them chart a sensible course for the future.
7. *I Know You're a Chemist, But What Do You Do?*; 1982, 8 pp.
   A brochure that features descriptions of first jobs and advice from newly employed B.S. chemists.
8. *List of Approved Schools*; published annually, 8 pp.
   A listing of institutions whose undergraduate chemistry programs are approved by the ACS Committee on Professional Training.
9. *The Chemical Industry and Careers for Chemists: An Annotated Bibliography;* 1988, 27 pp.
10. *Academic-Industrial Education News*; biannual newsletter, 8 pp.
11. *Career Options in Chemistry.*
    A compilation of reprints from an ongoing career series in the *pHilter,* the former ACS student affiliates newsletter (now *in Chemistry).* The articles discuss various chemistry careers, with each article focusing on a career specialty and featuring profiles of people working in that field.
12. *Chemical & Engineering News (C&EN),* published weekly.
    A weekly chemical newsmagazine and the official publication of the American Chemical Society. *C&EN* is designed to keep readers informed of the chemical world as it interrelates with the world at large. Each issue contains a combinationof short news stories, longer trend stories, and features on topics of broadinterest to the chemical world. All ACS members receive *C&EN.* Subscriptions for nonmembers are available from ACS.

13. *People Who Took Chemistry*; video (15 min.) and user's guide (52 pp) package. The video is useful as a tool for high school teachers to use with their students to encourage them to take high school chemistry and to consider a career in chemistry, and as a device to educate guidance counselors about the wide variety of careers in chemistry. The guide, written by an experienced high school chemistry teacher, provides the instructor with effective ways to use the video, including a variety of hands-on activities that can be used with students

Chemistry in Your Career was developed by James O. Schreck and Courtney Willis.

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

CROSS-REFERENCING OF MODULES

Part I

INTRODUCTION

The following is meant to provide the user of *ChemSource* with examples of how the 22 "core" modules can be used as references to one another, and particularly how the "enrichment" modules can be used to augment the "core" modules. Beneath each topic heading are listed the categories (*e.g., Laboratory Activity, Common Student Misconceptions*) and specific examples of related items that can be used to further illuminate the topics central to all introductory chemistry courses. Part I contains a list of the cross-referenced items for each of the core modules. Part II lists how each enrichment module can be used in conjunction with the core modules.

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SOLN 1: Arrhenius, p.26

Links and Connections

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BOND 3: "Mathematical or Mental Constructs such as Electron Clouds Represent Something Solid." p.35

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Links and Connections

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TEACHING CHEMISTRY TO PERSONS WITH DISABILITIES

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Teaching Chemistry to Persons with Disabilities

The inclusion of persons with disabilities in chemistry lectures and laboratories requires only a few changes. Consideration, kindness, and creativity comprise the most important ingredients to meet the needs of these students. The primary concern should be to provide the means for the student to acquire the concepts of chemistry rather than perform tasks incompatible with physical limitations. Given a means to acquire data and make observations, students should be able to interpret information, do calculations, draw conclusions, and assemble material into a laboratory report. The fewer the exceptions in procedure made for disabilities, the more successful the program will be, not only for the person with the limitation, but also for all the students in the course. If the teacher considers the abilities rather than the disabilities of students, a positive attitude will result, and the teacher will find it easier to cope with individual differences. All activities should be as normal as possible. There should be no exemption from laboratory investigation even when the person is so disabled he or she can only direct another to perform the experiment.

The curriculum should not change to “make things easier” for the student with a disability. With understanding and sensitivity by both teacher and student, the barriers to meaningful education can be overcome. Persons with disabilities can succeed in science. Itinerant or resource teachers assist by teaching students basic orientation and mobility skills. These specialists also locate and provide any special materials and equipment needed for academic instruction.

The number of persons with visual impairments who have majored in chemistry is small at present, but the numbers are steadily increasing. Other visually impaired students are choosing chemistry as a part of their general education or as a complement to other science courses. Persons with hearing and orthopedic impairments are well represented in all fields of chemistry, according to the Resource Directory of Scientists and Engineers with Disabilities. There is no real reason to discourage competent students with disabilities who wish to study chemistry.

TEACHER STUDENT CONSIDERATIONS

The Student Who Has a Visual Impairment

Preparation for a visually impaired student in chemistry class should begin following registration the previous spring. If the teacher or department head can be notified at that time, she or he can order supplies for delivery during the summer and make suitable preparations. If the student reads Braille, the chemistry text should be made available in Braille, not on tape. Also, it is important for the student to have a separate book on chemical calculations because much of this work will be presented on the chalkboard or with an overhead projector.

Before the first chemistry class, the instructor should have a conference with the student. A very important question is: “How do you plan to manage in the classroom and laboratory?” If the student has previously had science courses with lecture and laboratory, techniques for these situations will be familiar but should be reviewed carefully. Students may take lecture notes with a pencil, a Braille slate, laptop computer, or tape recorder. Some students may have already scheduled a friend in class to provide notes. For others, the teacher may have to find one or two good students willing to share notes *via* carbonless paper or a copy machine.

The teacher should ask the students where they would prefer to sit in the classroom. Some students prefer a place in the front in order to use what sight remains or to achieve better tape recorder reception. A few might wish to sit near a friend who guides them. Others prefer to remain as inconspicuous as possible.

The student who is visually impaired should have a tour of the classroom and laboratory before the first class. The student needs to investigate the contents of the cabinets and cupboards used by all class members. If the shelves are not labeled in Braille and/or large print, the student could be asked to help with the activity to insure that the Braille labels are not placed upside down.

Teachers who are careful about their techniques in lecturing and leading discussion probably will be helping all their students. The student with a visual impairment will be aided if the teacher describes orally all the material written on the chalkboard or shown on the overhead projector. When new scientific terms are used in the lecture, it is helpful if the teacher spells an occasional word. Persons who never see a word and who use grade two or three Braille, which uses symbols for groups of letters, will have difficulty with the spelling of unusual words. Diagrams should be described in a consistent manner starting at the top and going around clockwise. The words *this* and *that* should be avoided unless used with the noun to which they are referring. Naming a substance is more informative than the word *it*. When there is a change in temperature, pressure, mass, or some other characteristic, it should be described in precise numerical terms instead of saying it goes from *here* to *there*.

The teacher should follow a consistent and specific procedure for giving assignments, which might be distributed in printed form or written on the chalkboard. The student will find a way of having it read so he or she can Braille it or put it on tape. Likewise, one must remember that a chemistry text in Braille will consist of ten or twelve volumes for a single book. The teacher should specify which chapter he or she wishes brought to the class, thus lightening the student's load.

The teacher will subtly help the student if roll is taken orally. This technique makes it possible for the student who is visually impaired to learn to recognize the other students' voices and know who is in the classroom that day.

When demonstrations are to be presented to the class, it will be helpful if the teacher will give the students with visual impairments the opportunity to go over the equipment manually so they will know the positions and functions of the parts. Such activities also can teach the other students some techniques for working with persons who have limited sight.

It is important to select the most suitable laboratory partner for a student with visual impairment. A good student, who is also a long-time friend, would be the ideal choice. The partner must be someone who will expect and allow the person who is visually impaired to do his or her share of the work. Persons with disabilities are less likely to have accidents in the laboratory than are students without limitations. Students with visual impairments are quite accustomed to taking care of themselves and being very careful with everything they handle.

One must remember that of those persons who are designated legally blind, only ten to twenty percent will be totally without vision, or truly blind. The others will vary from those who can see stop lights and large letters with the aid of a monocular clipped to their glasses to those who, with strong glasses, are able to read regular print if the paper is held at the correct distance. A few of the latter manage to acquire an education without disclosing their disability. The group between these two extremes employs a variety of magnifying devices for which light requirements will vary. The itinerant teacher for the visually impaired, the resource specialist, or the parent can undoubtedly advise the teacher on this subject.

REPORTS AND TESTS

If the laboratory manual contains blanks and charts to be filled in as the report, the manual should be brailled before the opening of school. With both Braille and print, adequate space must be allowed for locating the answers. Label tape is available for both modes. The tape is then placed in the appropriate space.

By the time visually impaired students reach high school, they should be skilled in the use of the typewriter, computer or VersaBraille. Blind students who read a grade two (average) or grade three (advanced) Braille will be encountering symbols for groups of letters. As a result of this particular skill, their spelling is sometimes phonetic. The teacher needs to realize such problems.

High school chemistry students often work in pairs. However, that is no excuse for a joint or duplicated laboratory report. The data may be the same in two reports, but the observations, calculations, evaluations, and conclusions should be those of each individual. If drawings are expected for a report, the student should make a raised line drawing which is labeled in both Braille and print. If a graph is needed, the student should be able to prepare one on brailled graph paper or on a computer. Lines on Braille graph paper may be made with a wax pencil or tracing wheel. Chemistry problems for a test or a report can be worked on a talking calculator used with earphones or with an abacus. A raised line drawing may be sufficient for balancing equations. (At the college level students may have the use of an Optacan, a scanning device and a computer with voice input and output to perform such tasks more efficiently.)

Major tests should be brailled for those who read Braille. Short quizzes should be produced with the brailled labeler. All other tests should be taped, not only for persons with visual impairments, but also for those with disabilities which limit normal reading skills. A tape recorder can be used for recording answers for those who cannot type or write. Computers with speech capabilities are also convenient for use in testing. All students should take the same test at the same time and preferably in the same place. A typewriter should be available for the use of students who cannot write script and when more than a one-word or letter answer is anticipated. Most students with limited sight write well enough to place a letter or one word on a sheet of paper. Special enlarged forms for multiple choice test questions are not as practical as allowing the student to choose his or her spacing on a plain sheet of paper.

Having all students come to class at the same times saves the student with disabilities peer criticism. Also, keeping the students with limitations to a schedule helps them to adhere to their own schedule and extra-curricular activities. Students with handicaps wish to do the same as the other students in every way possible.

In general, students with visual impairments should expect to be graded on the same scale for laboratory work and tests as is the rest of the class. If the major criterion is to show evidence of acquiring the concepts of chemistry, collecting, interpreting, and evaluating data plus drawing valid conclusions, then he or she should be graded accordingly. If the techniques and instruments a student with a disability uses, because of his or her disability, are less accurate than equipment used by the rest of the class, then there should be allowance for this in grading performance. It may be necessary for one who reads large print or Braille to be given more time. A student who is hearing impaired can see and write but often reads below grade level. Likewise, students with orthopedic or neurological disabilities may take longer to take a test because of poor manual dexterity. Assistance from an aide in reading tests and writing answers for a student should be a last resort. Most students with physical limitations will appreciate consideration and try to compete with peers. There is little tendency to take advantage of their situation.

SPECIAL EQUIPMENT

The array of new instruments which has become available in recent years has made chemistry a more popular course with students who are visually impaired. Since greater independence is now attainable in the laboratory, lecture hall, and office, more persons can consider chemistry as a vocation. A visually impaired person should not postpone high school chemistry because the school lacks these sophisticated new tools. The concepts of chemistry are available to all who are willing to listen, read, and learn. Today's student can look forward to the use of a well-equipped laboratory in college that will give him or her more independence.

At the high school level, the itinerant teacher of the visually impaired or the State Department of Special Educational Needs will often lend the special equipment needed for a visually impaired student to be able to participate as fully as possible in the chemistry laboratory. Also, certain simple devices can be purchased or created for the benefit of student or teacher.

Useful devices include the **Braille Labeler**, an instrument that embosses Braille on vinyl or magnetic tape. The rim of this labeler dial contains both Braille symbols and letters as well as mathematical numbers and signs, so that the sighted person who does not read Braille is also able to produce labels. This device can be used for writing short quizzes, grades, and short comments for the teacher to affix to tests. An optional dial for making large print is helpful in labeling bottles, containers, supply cabinets, drawers, *etc.* in the laboratory.

In addition, well-labeled containers, and bottles with different surface types and colors are available. For example, one bottle may have a strip of sandpaper around the middle that can serve as a warning. Ribbed and smooth containers can also be used for a variety of purposes.

Braille labels should be placed on all shelves in the laboratory to serve as a reminder to the student with sight to replace things where they belong. In the absence of a suitable labeler, a student with visual impairment can make Braille labels by using a sheet of self-adhesive labels in Braille slate. A sighted laboratory assistant can spell out the words to be used and affix the labels to the correct bottles.

A **peg board** with golf tees to make the points that are connected by string, yarn, and cord to show the lines makes a practical graph and an excellent demonstration for all the members of the class.

Braille writers are usually the property of the blind students. These are larger than a Braille slate and are used with a sheet of Braille paper. They enable the person to print Braille more rapidly than on a Braille slate. The older noisy models can be reserved for use in the laboratory. A new, quiet model is now available. Its site of use should be determined by the teacher.

Braille graph paper is usually supplied by the student. If it has been forgotten, the student can easily make it on the Braille writer. The student uses a wax pencil to make a mark on the graph that is *seen* with the fingers and the eyes. A bar graph can be constructed by using wax crayons. A tracing wheel from the sewing department makes a dotted, raised line that also can be *seen* with fingers and eyes.

Organic Chemistry Molecular Model Kits are an excellent educational tool for all students. Synthesis and decomposition of organic molecules, when done with these kits, can be more memorable. The student who is visually impaired recognizes **H**, **O**, **N**, and **C** by the number of holes for bonding. For **S** and **Cl** tiny brads can be added as markers. Additional wooden shapes are added to represent a benzene ring and other common groups in order to save time in construction. (These additions are not available commercially but are easily home-made or produced in the school wood shop.)

A student who is blind should have a **Braille Ruler** or a plastic ruler with raised markings. The lines on the standard meter stick are indented sufficiently to be counted by tactile means. Braille numbers made on label tape could be added to make a Braille meter stick.

The **Sewell Board** consists of a one-fourth inch thick rubber mat on which Braille paper or aluminum foil is placed for making raised line drawings. Similar boards made at home are quite suitable. Several layers of wool cloth tacked to a 10" by 14" board will do nicely as will several layers of rubber cut from an old inner tube.

Raised line drawings are frequently needed to explain the positioning of equipment for a chemical reaction or for copying sketches shown on an overhead projector. These drawings require no particular artistic skill or ability. They are similar to quick sketches on paper except that they are made on Braille paper with a stylus or tracing wheel. If a laboratory report requires a drawing, the student with visual impairment attaches one of these sketches to his or her report.

Students with marginal vision often have difficulty seeing material presented on the overhead projector. Transparencies may be copied and enlarged on a copy machine. Drawings thus produced may be traced with a felt marking pen to produce a **Large Print Drawing.**

Hi Marks Labeling Substance is a three-dimensional fluorescent plastic marker that comes in a tube. It is bright orange and creates raised markings on paper, wood or metal. It dries slowly.

The ink tracing produced by a gas chromatograph and other instruments is often too small to be observed tactilely even if the line is raised. A photocopy machine enlarger can produce a copy whose lines can be raised with a tracing wheel.

Liquid Level Indicators available from Science Products, are plug-in accessories for the **Audicator** (used with the light probe). The 4" and 12" steel probes clip to the side of the container and can be set to the level desired. A two pronged probe is a special liquid level indicator. Students should be discouraged from using a finger on the edge of a container to locate a liquid. Clip Leads for the Audicator can check bulbs, switches, fuses, circuits and electrical cords.

Students with visual impairments often need assistance in learning to balance equations and do chemistry problems. The problem of communication in printed form may be solved with a **Magnetic ABC Board** in Braille and large print. The set of two-inch letters and numbers make a good common denominator. A **Braille Scrabble© Board** with the addition of homemade numbers also can serve. Pieces of wood with painted numbers need only have the numbers outlined with dried white glue (such as Elmer's Glue-All©).

Brailled or taped supplementary texts and journals in chemistry are rare at present. Such materials must be presented by a reader. In some communities, volunteers will Braille or tape material on request. A few students own or have access to an Optacon, which is manufactured by Sensory Access Foundation and is a general purpose reading device. It consists of a miniature television camera guided by a finger on one hand while the tactile display is felt by the forefinger of the other hand. It is adapted to read almost any printed material. With a special adapter the Optacon can be used to read meters of laboratory instruments. This device is too expensive to be part of the equipment in most high school chemistry laboratories.

The **Light Sensor** (light probe) is an important laboratory instrument for the visually impaired person. This device has a battery-operated photoelectric cell connected to a sounding mechanism or audicator. The greater the intensity of the light, the higher the pitch of the sound produced. It may be used with an earphone so others in the room will not be disturbed.

Among the many uses of the light probe are:

1. locating the meniscus of a clear or colored liquid in a graduated cylinder, volumetric flask, or a buret. The raised or etched markings on the outside of the container are counted with a fingernail or a thin plastic card to determine the quantity. The accuracy of this type of reading is not quite as good as a visual one, but for providing independence in the laboratory and teaching the principles of various chemical experiments, it is an asset. Grading on laboratory skills should be related to the accuracy of instrumentation.
2. sensing the color change as in the endpoint of a titration. It can also sense the color changes in qualitative analysis.

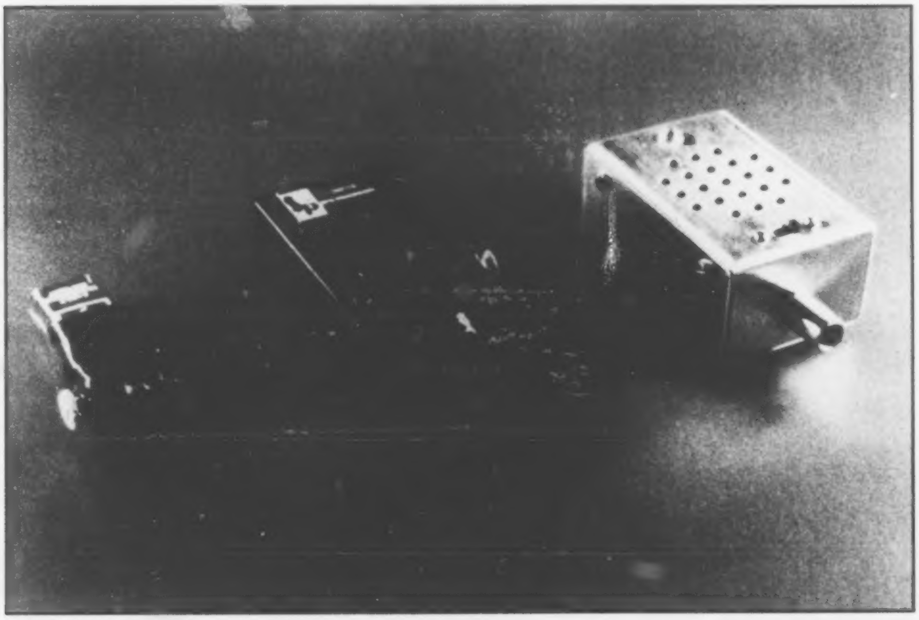
3. matching colors of solutions.
4. sensing the difference between clear and cloudy solutions, thus indicating the formation of a precipitate.
5. sensing the formation of gas bubbles.
6. locating the colored spots on paper and the penciled base line for chromatography.
7. finding the margins on a typewritten page.
8. locating the needle of the dial on various laboratory instruments.

   A second scale with Braille markings can be made for the device. At home it will tell if socks match, if clothing is plaid, striped, or plain and whether an apple is a McIntosh or a Golden Delicious.

Three light sensors are shown in Figure 1. The instrument in the center is produced by Science Products. Designed by a college instructor who is blind, its sensor can be extended on a cord while the audicator remains in a pocket. It is an excellent, sensitive unit. A small probe for this sensor designed by Smith-Kettlewell is available from several sources.

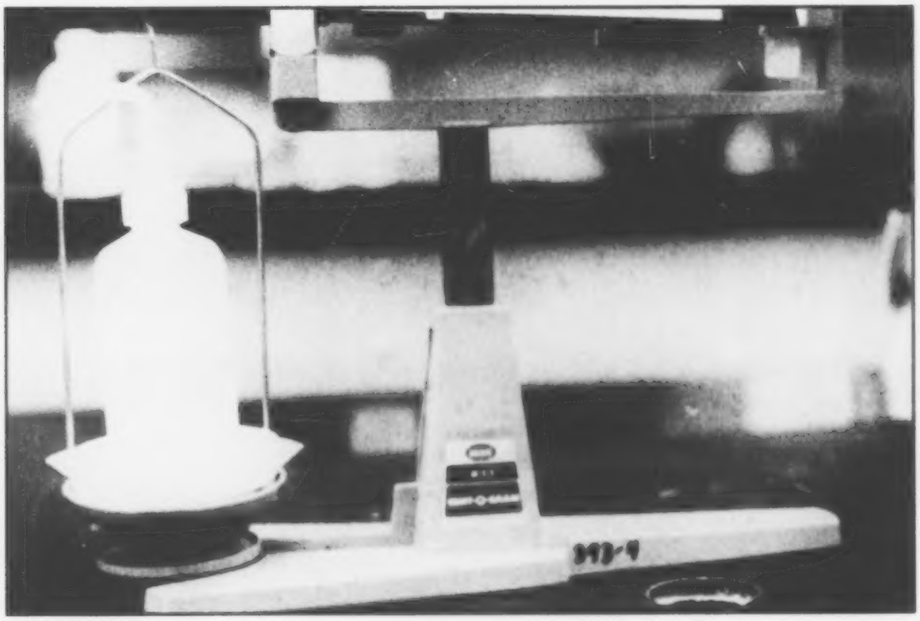
The unit on the left in Figure 1 is from American Printing House for the Blind. It is not precise enough for determining such things as a meniscus, but it is durable and suitable for many uses in the laboratory and at home.

Figure 1. Light sensors



The Weight Buret (Figure 2), described by Seils, *et al.*, was designed for accuracy. It consists of a plastic wash bottle to which is attached a pipet with a fine point. The wash bottle is filled with titrating solution and weighed before and after the titration. The students with limited sight determine the endpoint by the change in sound of the light sensor that has been firmly placed in a clamp on the ring stand. Accuracy can be enhanced by having very even lighting such as that given by a fluorescent light covered with a piece of ground glass. The beaker with the solution to be titrated is placed on a magnetic stirrer. Hand stirring the solution might interfere with the beam of light to the light sensor that is located a few inches from the beaker. The pre-weighed weight buret is squeezed gently to allow the solution to go into the reaction a drop at a time. There have been reports that claimed greater accuracy in determining the endpoint with the light sensor than with natural vision. This technique is also suitable for persons with orthopedic disabilities and limited manual dexterity.

Figure 2. Weight buret.

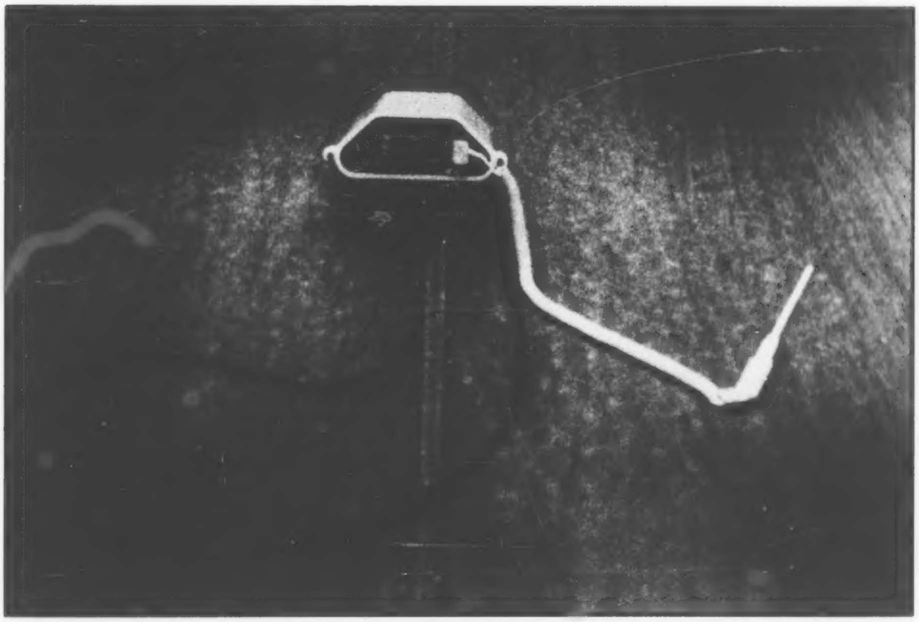


SPECIAL INSTRUMENTS

Talking Calculators were the first instruments to have voice output. These calculators are an asset for calculating chemistry problems. When used with an earphone, they are not audible to others in the room, so they serve nicely during tests without sharing information. If the student with visual impairment does not own his or her own talking calculator, there should be one available at the school or from a state source for use in mathematics and chemistry classes.

In a survey of chemistry and biology experiments, it was discovered that thermometers were used more often than any other instrument. The **Talking Thermometer, ThermoVoice™**, (Figure 3) was designed by the American Foundation for the Blind to serve at home and in the laboratory. Thermovoice has a plastic coated stainless steel probe that registers in the range of -10°C to 110 °C. Only a few seconds are necessary for the probe to stabilize after being immersed in the material to be tested. Temperatures registered at below zero are reported by "minus." Another probe, registering from 27 °C to 210 °C, can be used in organic chemistry or in the kitchen. If the wrong probe has been selected for the temperature of the material the Thermovoice says "over, over." The precision of Thermovoice is ±0.2 °C.

Figure 3. ThermoVoice.



Thermovoice can be worn on a cord around the neck, thus allowing the hands to remain free. It can also be operated while using an earphone, and the latest Thermovoice model operates on a nine-volt battery. Chemists with sight also appreciate using the talking thermometer when their goggles or glasses are foggy, when the thermometer is located above eye level, and for many other situations.

Several types of Talking Computers are now available, and a computer with a Braille readout is also on the market. These computers are sometimes available at the school computer center or at the local Society for the Blind. Large print models are also available.

Most students with visual impairment will have their own Braille or talking watch. Talking Clocks are available that speak every hour or on demand. These clocks may include an alarm. In the laboratory, the big wall clock that has second and minute hands and is designed to be operated manually will be serviceable. White glue can be added to the lines to indicate divisions and to the numbers to make tactile reading possible. Also, a kitchen timer that has raised or indented lines can serve for timing experiments.

Several manufacturers market a Talking Blood Glucose Meter for monitoring a patient's blood glucose. This meter would not be appropriate for the chemistry laboratory, but it might make an interesting demonstration for a class.

Students with visual impairments become adept at weighing. The Ohaus Double and Triple Beam Balances lend themselves to proficient use by those who are sight impaired. The weights on the beams slide into notches that are readily counted by touch to determine the value. A piece of tape placed at the zero point serves as a guide to tell when the pointer is centered on zero. For the Cent-o-Gram Balance with a quadruple beam, the slide with no notches can be calibrated with the student's Braille ruler such that 1 cm equals so many milligrams. More sensitive balances may be read by a light sensor to determine when the pointer is centered.

The newest addition to the family of talking instruments is the Port-O-Gram Electronic Ohaus Balance to which the American Foundation for the Blind has added a voice box. Any present balance of this type in the laboratory can have a voice box added if the instrument has a serial port. The teacher can anticipate additional voice-output laboratory instruments in the near future.

Modifications of the Bausch and Lomb Spectrophotometer for use by students with visual impairment were made by Lunney, Morrison and Hartman. These modifications consist of a recorder output that intercepts the electrical signal to the meter and a tactile dial for setting the wave length. Some new instruments have recorder outputs as standard equipment.

When sophisticated accessories are lacking, a light sensor can locate the needle on the dial of the Bausch & Lomb Spectronic 20. Braille dots can be put on the dial to give numerical value to the reading. Total independence is available here at the cost of diminished accuracy. When accuracy is essential, the readings need to be done by a person with sight.

Rawls describes a standard spectrometer adapted by Cochin and Skawinski, for the use of persons with visual impairment. In place of an eye piece, a light sensor is mounted on a Braille micrometer. A prism produces the spectrum that is scanned by slowly adjusting the micrometer. Using this method, the operator can concentrate upon one spectral line at a time. The pitch of the accompanying oscillator is directly proportional to spectral line brightness.

The Kurzweil Reading Machine provides personal, direct access to printed and typewritten information. The user places the material to be read face down on the glass surface of a scanner. A separate compact control unit is then activated by the user. The scanning mechanism automatically locates the first line of the text and begins scanning the page. Within a few seconds an electronic voice is heard reading the material. The user has at his or her command controls that can select 33 items such as speed, tone, and recall of a few lines. This instrument likely to be found in college and large public libraries.

VersaBraille II+ (Sensory Access Foundation) is a portable laptop computer that visually impaired students can use to take class notes, complete exams, do homework and write papers. It also can be used as a console to interface with a personal computer. The keyboard of the VersaBraille II+ has six keys that corresponds to the six dots of Braille. The student uses the six keys to type grade two Braille rather than English orthography. When students want to check their work, they can read a Braille display located on the top of VersaBraille II+.

VersaBraille can be connected to computers and other devices. It is possible for a student to prepare an assignment on the VersaBraille and then transfer it to a computer that uses a talking word processing program such as Braille-Edit Xpress (BEX). The text can be edited and then printed out in a standard print format for the student's teacher to read. A Braille printer can be used to provide the student with a Braille copy of the material. Some specialized printers produce pages with text in both Braille and print formats so that both blind and sighted readers can work from the same copy. The VersaBraille II+ system allows a blind student to participate in mainstream classes without a note taker. It is also a valuable tool for blind adults in the workplace.

Portable text-to-speech synthesizers are available with good voice quality, unlimited vocabulary, and automatic inflection control. These synthesizers are designed for those who have good hand and finger dexterity. Available in easily transportable units they will speak what the user has typed on command.

SMALL-SCALE CHEMISTRY

Small-scale laboratory experiments in chemistry may be an advantage or a disadvantage, depending upon the nature of the student's disability. Visually impaired students benefit from enlargement of the scale of materials used, since many specialized measuring devices and sensors require a sufficient quantity of reagent to activate the sensing device. On the other hand, those who are orthopedically impaired benefit from a diminished scale of materials since smaller pieces of laboratory equipment can be more easily handled. Teachers will have to use their best judgement in such cases, but the following points may help when adapting small-scale experiments for the visually impaired:

1. Amounts that can be weighed or measured with a syringe are more practicable than counting drops. Drops are too quiet to count by listening to their fall.
2. Beral™ pipets are not suitable since sight is necessary to control the quantity of fluid in the bulb.
3. Disposable plastic syringes that can deliver small amounts precisely are practical for the visually impaired. The syringe barrel can be notched for tactile reading.
4. If small-scale well plates are used in conjunction with a light sensor, (a) samples should be placed in alternate rows and columns so that material in adjacent wells does not interfere with the light sensor readings; (b) the sensor must be placed directly over the well to be sure that the light beam is not at an angle; (c) the well plate should be placed on a contrasting background (*e.g.*, black for white precipitates) to maximize light sensor sensitivity.

CONSIDERATIONS FOR THE STUDENT WHO IS HEARING IMPAIRED

Meeting the needs of students who are hearing disabled in the lecture room and laboratory is quite different from addressing the needs of students who are visually or orthopedically disabled. For students with impaired hearing, everything revolves around means for good communication by student, classmates and teacher. It is important to have a conference with the student, preferably before school starts, to discuss how he or she plans to manage and what needs to be done for personal amplification.

Few persons who are profoundly deaf are sufficiently adept at lip reading to obtain the information in a lecture and to participate in class discussion. Students with hearing disabilities often wear hearing aids, but this does not mean that they hear everything that goes on in the classroom. Hearing aids do not give "normal hearing." Sounds are louder but not necessarily clearer. Background noise also is amplified. Some sound, even if not perfectly clear, will give cues to aid lip readers. There is a wide variation in general skills between students who neither hear, lipread, nor speak and those who can discern some sound.

The choice of seating in the classroom should be discussed at the early conference. A seat at the front of the room is usually chosen for better sound reception and a better view. Some students who have a moderate loss may be using an assistive listening device (ALD) in the classroom. This device could be an induction loop, FM, or infra-red system and might require a clip-on microphone for the instructor, who should speak in a normal manner. The speaker should also avoid standing in front of a light source as it makes the face and lips difficult to see.

A sign language interpreter may be available during lectures. This person is usually seated at the front and toward the side of the classroom to enable the student to see the interpreter, the teacher, the blackboard and the screen of the overhead projector. The use of an interpreter enables the student to participate in class discussions and to know what is being said by other students as well as the teacher. If interpreters are not familiar with chemistry, the teacher should provide a vocabulary list beforehand so that the interpreter and student can agree on abbreviations. This provision is particularly important when signing compounds, which take a long time to fingerspell. When speaking to the student, the teacher should look at and address the student, not the interpreter. Students who are hearing disabled need to be called on during discussions. If there is a signer present, this is easy to do, and the student can then be drawn into discussions with other students. If there is no interpreter, the teacher should use appropriate eye contact or gestures to call on the student. A note pad should be readily available for times when other methods fail.

How deaf students obtain class notes is a matter that must be clarified early in the year. A student who lipreads, or watches a sign language interpreter, cannot take written notes or read written material at the same time. Students old enough to enroll in chemistry should take the responsibility of obtaining their own notetakers. However, because of problems in communication, they may have had difficulty interacting with other students socially and be hesitant to ask for someone to take notes for them. The teacher or counselor should find one or more volunteers willing to be notetakers for the student who is hearing disabled. Some teachers are willing to supply copies of their own notes to the student. Whenever possible, students should be provided with printed class outlines, reading assignments, and scripts of audio-visual material. Likewise, test schedules, grading policies, and requirements for term papers or projects should be given in written form. Copies of transparencies shown on the overhead projector also are helpful.

Overhead projector technology in a high school class serves all students except the visually disabled. Slides should be shown with printed labels. When pointing to something on an overhead, allow the student time to look at the interpreter (or your lips) and then back at the projection. This will also give other students valuable processing time. As with sight impaired students, avoid using words such as "this" and "that" when pointing to an overhead. When films and videos are used, they should be close-captioned, if possible. When new vocabulary occurs in lecture or on video, it benefits the whole class to write the new words on the chalkboard or overhead. Signing is not a direct translation of English. It expresses everything in the positive; be patient with a signer who asks you to repeat something, especially if it is technical. Signers often must wait until the end of your sentence to rephrase something in the positive. Likewise, avoid asking multiple choice test questions such as "Which of the following is NOT . . . ." Also be cognizant of the fact that we normally do not communicate in complete sentences; an attempt should be made to make sentences as coherent and complete as possible for the hearing impaired student.

In general, impaired hearing has little effect on the ability of a person to work in a chemical laboratory. Visual warnings such as flashing lights should be installed as a fire alarm. These devices can also be wired to the school's bell system, but students usually know when class has ended. Instruments should be equipped with lights to indicate on/off status, where possible. Emergency lighting should be available in case of power failure. When the room is darkened to show movies or videos, there should be a small amount of light in the room so persons who are hearing disabled can orient themselves. The best situation in the laboratory is with a compatible laboratory partner who shares the work and finds individualized techniques for communication. Lipreading is difficult in the laboratory since persons tend to be more mobile. Teachers should check occasionally to see that the laboratory partner is passing on announcements to the student who is deaf, and check with the student to see if the notetaking is satisfactory and if there are any other ways in which the student can be helped.

There are many new devices to aid communication for persons who are speech and hearing disabled. These instruments are usually quite expensive and not readily available in all areas. The moisture and fume laden atmosphere in a laboratory makes these devices more suitable for the classroom than the laboratory. New devices are available which will voice a message typed into it, and others will print words spoken to it. Voice input and output can be interfaced to computers. Any device that prints out spoken words improves the reading level of students who are deaf.

We are in a period of transition. It is just a matter of time and economics before communication skills can be vastly improved for the students who are hearing disabled. Many persons who are hearing disabled have been successful as chemists at all levels. It is important not to discourage students who are hearing impaired from pursuing a vocation that has offered satisfaction for persons with similar disabilities.

Classmates of students who are hearing impaired often wish to learn sign language. Some schools have classes or clubs available for this purpose. Others are available for those who wish to learn to Braille so they can write notes to/for their friends who are visually disabled.

STUDENTS WITH ORTHOPEDIC DISABILITIES

Persons who are orthopedically disabled are well represented in the Resource Directory of Scientists and Engineers with Disabilities, published by the American Association for the Advancement of Science. High school chemistry departments can expect to find students with a wide range of inherited and acquired disabilities in their classrooms and laboratories.

Preparation for students with orthopedic disabilities in the chemistry classroom and laboratory should include departmental policy decisions before the need arises. Students with permanent disabilities can be recognized at registration time. On the other hand, the presence of a student with a temporary disability will be unpredictable. The teacher must be prepared for both.

Long term planning might include provision for adjustable height laboratory benches, adjustable height storage units, and special equipment work spaces. It would be ideal to have several adjustable height counters for all classes. When students in one class varied in height from 55 inches to 80 inches, "one size fits all" was a myth with reference to a standard 36-inch high chemistry laboratory bench. Single action lever controls or blade-type handles rather than knobs are needed for students who lack good manual dexterity or have difficulty reaching across a bench. Flexible connections for electric, water, and gas lines plus a means for disposal of liquids and solids must be provided for students who have limited reach (*i.e.* those in wheelchairs or with impaired arms). Several companies that manufacture laboratory work stations have units designed for wheelchair users. These also have specific hoods available. The unit may even be mounted on casters to provide mobility and availability in more than one laboratory. The standard hood in a chemistry laboratory is a place best avoided by persons in wheelchairs. If a new laboratory is being designed, aisle width of 42 to 48 inches should be considered to accommodate wheelchairs. The eyewash fountain in the chemistry laboratory should be checked for correct height for use from a wheelchair. Student work stations should be located near an exit to insure easy egress in case of emergency.

When commercial units are not available, a somewhat improved laboratory station for students in wheelchairs can be improvised. A work bench can be assembled from a 30-inch high biology table or business department table.

Laboratories can offer some special problems in mobility to those dependent upon wheelchairs, crutches, or walkers for movement. For example, while laboratory aisles may actually be wide enough for the passage of a wheelchair or a turn-around, when class is in session such movement may be difficult. A student of average height will have trouble reaching across a laboratory bench from a wheelchair. It is sometimes painful and always tiring to stand on crutches for long periods of time. Carrying chemicals while using crutches or operating a wheelchair poses definite potential hazard to other students in the laboratory.

Larsen developed a platform chair in order to be able to work at a chemistry laboratory bench of standard height and to be able to move about the room. He used an oak desk chair securely fastened to a 20-inch x 24-inch platform supported by heavy duty casters at the corners. This chair occupied less than half the floor space needed for a wheelchair. The chair had a tray beneath it in which to carry chemicals and supplies to the work station. R. Tombaugh further modified the platform chair by using a folding metal chair held to the platform by metal brackets and wing nuts. This chair has the advantage of taking up little space when folded and disassembled for storage in the stockroom. The chair is propelled by using adjacent objects on which to push. In the confines of a single room, there is no need for brakes.

All students with disabilities should have a conference with instructors and laboratory aides before starting the course. The student should be asked how he or she expects to manage in the classroom and laboratory. An opportunity should be given for the student to try a work station to see if the arrangement in the laboratory is convenient. For students with manual disabilities, it is important to determine their planned method of note taking. They may need note takers since tape recordings of problem solving and balancing equations are not satisfactory. A student volunteer can make a carbon copy as he or she takes notes. Some teachers provide copies of their lecture notes. Some students may need a reader for tests who will also record their answers. This may be done in another room but should be at the same time as when the rest of the class is taking the test, thus minimizing criticism of favored status by peers.

For students who have limited hand movement, it is advantageous to have smaller items such as those found in small-scale chemistry settings. Students who cannot open their hands sufficiently to pick up a 100-mL beaker, may be able to manage 7-mL reagent bottles, 10-mL graduated cylinders and spot plates. The use of 24-well plates is quite suitable. Plastic syringes (or Beral™ pipets) are often better for measuring liquids than are capillary tubes when students lack fine motor control in their hands.

Small-scale chemistry is a distinct advantage for persons in a wheelchair or who work in a seated position. Reagent bottles can be small and each station can be equipped with its own set of chemicals. It is hazardous for a student in a wheelchair or on crutches to move through the laboratory to a central supply area and return to the laboratory station with chemicals.

Talking instruments such as ThermoVoice™ are also useful for persons working in a seated position. It allows the taking of temperatures with a probe that is not at eye level. Special racks may be used in addition to ring stands to bring equipment within reach of the person who is seated when performing experiments.

For a student with very short arms, there is a possibility that his or her face may be too close to the chemical reaction. A clear plastic face mask is better than safety goggles in this case.

For students who are too immobile to perform any laboratory procedures, being a member of a laboratory team gives them a chance to participate. They can help make decisions for procedures and can make observations of the results. They can do their calculations and draw conclusions independently. When they are in the laboratory, they will need aprons and goggles or a face shield just like anyone else.

Chemistry experiments done entirely *via* computer are becoming more common.

These experiments are advantageous for students who are orthopedically or neurologically disabled since their need for movement is diminished. However, the future chemist must have adequate hands-on experience in the laboratory.

LABORATORY SAFETY

General rules for safety in the laboratory are given elsewhere in *SourceBook*. These additional rules are oriented particularly toward students with physical disabilities.

1. Give the student with impaired vision an opportunity to become familiar with the laboratory before the first session. The student can then participate in the safety-oriented program with little trouble and will already know the locations of exits, showers, and extinguishers.
2. A person who is visually impaired should have a laboratory partner.
3. Discuss and resolve individual limitations with the student. Can the student read labels? Are the labels large enough? Does he or she feel any operations are too hazardous.
4. Assign the student with impaired mobility to a laboratory station on an outside aisle and close to an accessible exit.
5. Students, including those with impaired vision or poor manual coordination, are urged to wear rubber gloves when working with harsh chemicals or those readily absorbed by the skin. Disposable, lightweight gloves are available that will permit the student to manipulate equipment.
6. Keep all aisles free of obstructions. Crutches also should be kept in a secure place where they are not apt to fall into the aisle.
7. All students should wear plastic or rubber aprons when working with chemicals. Students in wheelchairs should be advised of the importance of protecting their laps with a heavy rubber apron.
8. Persons with extremely short arms may need a whole face shield instead of safety goggles.
9. Accessible and usable eye washes should be located near the work station of a student who is disabled.
10. Lightweight fire extinguishers should be provided for students with limited mobility.
11. Laboratory alarm systems should be visual, with flashing lights for students with hearing impairments.
12. Carrying chemicals while walking with crutches should be forbidden.
13. Some techniques are usually taught to a class by demonstration. These demonstrations should be checked to see if the laboratory partner has correctly taught the student who is visually impaired. The correct procedure for removing a stopper from a bottle of chemicals, for inserting glass tubing in a rubber stopper, and other items involving assembly of chemical and electrical equipment must be verified.
14. Persons with limited sight are accustomed to greater use of their olfactory sense. They should be taught the method of wafting odor, and should be warned of the toxic nature of many chemicals and the dangers in smelling the contents of bottles of chemicals in the usual manner.

15. The use of electric hot plates in place of burners should be considered for students who are visually disabled or for those who work in a seated position.

Experience with one's own disability has proven to the individual that foresight is indispensable. Since careful preparation is the keystone of laboratory safety, the person with a disability, by virtue of a lifestyle that requires planning, is actually better prepared, and hence often poses less of a safety threat than do students who do not have physical disabilities.

References

This module was developed by Dorothy Tombaugh, with assistance by T. Kucera of the ACS Committee on Chemists with Disabilities.

NOTE: The March, 1981 issue of the *Journal of Chemical Education* contains a series of articles on teaching chemistry to the physically handicapped. While these articles are referenced below, it would be to the teacher's advantage to obtain this issue from the Publications Coordinator, Kenneth Emerson, *Journal of Chemical Education,* Subscription and Book Order Department, 1991 Northampton St., Easton, PA 18042.

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SELECTED RESOURCES

* Starred entries are sources of supplies and laboratory equipment for persons who are visually disabled

American Association for Advancement of Science (AAAS), Project on Science Technology and Disability, 1333 H Street, NW, #1011A, Washington, D.C. 20005.

American Chemical Society (ACS), 1155-16th Street, NW, Washington, D.C. 20036. Committee on Chemists with Disabililities.

* American Foundation for the Blind, 15 West 16th Street, New York, NY 10011.

* American Printing House for the Blind, 1839 Frankford Avenue, Louisville, KY 40206.

Apple Computer, Inc., Worldwide Disability Solutions Group and National Special Education Alliance, 20525 Mariani Avenue, Mailstop 36SE, Cupertino, CA 95014. (408) 974-7910. Information and technical assistance is available regarding computer technologies available for use by persons with disabilities, including: "Solutions," a Hypercard stack of computer resources for individuals with disabilities, and "Connections," a guide to computer resources which may be used with Apple computers for persons with disabilities. Also see "When Versa Goes Apple for a Blind Chemist," by Cartier, Jean-Pierre and Jones, Peter, *Journal of Chemical Education, 65,* 525-37 (1988).

* Arkenstone, Inc., 1390 Borregas Avenue, Sunnyvale, CA 94089, (800) 444-4443, (408) 745-6739 (FAXDisabled Children's Computer Group, 2095 Rose Street, Berkeley, CA 94709. (510) 841-3224. This non-profit organization is a provider of a range of reading machines for the blind and visually impaired as well as those with reading impairments. Also see Kurzweil under Xerox, p. 22.

* Braille Book Bank, Braille Materials Production Committee, National Braille Association, Inc. (NBA), 1290 University Avenue, Rochester, NY 14607. (716) 437-0900, (716) 473-4374 (Fax). This committee coordinates all the duplicating and transcribing services for NBA, for college texts, mathematics and scientific tables. Catalogues of current materials in Textbooks and Technical Tables are available. Braille transcribing services are also available.

* Braille Technical Tables Bank, National Braille Association, Inc., c/o Mrs. James O. Keene, 31610 Evergreen Road, Birminghan, MI 48009. The Association supplies thermoform copies of mathematics and scientific tables.

* Enabling Technologies, 3102 Southeast Jay Street, Stuart, FL 33497. (407) 283-4817. Puts out an IBM based Braille graphics package, ET Graphics, which can import images from paint programs using formats like PCX and TIF.

Foundation for Science and the Handicapped, Inc., 154 Juliet Court, Clarendon Hills, IL 60514.

IBM Independence Series Information Center, Building 5, 3rd Floor, P.O. Box 1328, Boca Raton, FL 33429. (800) 426-4832 (Voice) or (800) 426-4833 (TDD).

* Lighthouse for the Blind, 1155 Mission Street, San Francisco, CA 94103. (415) 431-1481. The organization supplies a light probe (emits a tone which increases in pitch with changes in light intensity - often part of readout devices) as well as a talking calculator and large print and bold line papers.

* L.S. & S. Group, Inc., P.O. Box 673, Northbrook, IL 60065. (708) 498-9777. Provides a catalogue for instruments and devices for the visually impaired and those with communication problems.

Macrolab, Center for the Handicapped, New Jersey Institute of Technology, Newark, NJ 07102. Drs. Ira Cochin and Harry Herman.

* Raised Dot Computing, Inc. (RDC). 408 South Baldwin, Madison, WI 53703. (608) 257-9595. Publishes a bimonthly newsletter for their products and other helpful hardware, software ideas and comments for the visually impaired. RDC has a broad graphics package called pixCELLS for the Apple II.

* Recording for the Blind (RFB). 20 Roszel Road, Princeton, NJ 08540. (800) 221-4792. This organization provides cassette tapes of educational textbooks (from K to postsecondary levels). 78,000 titles are currently available; other titles are recorded upon request. Electronic text versions of some technical books including computer manuals are available for a nominal fee. After a $25 onetime fee this service is free for the blind, the visually impaired and for those with a documented learning disability; the materials are on loan. There is a free 'E-Kit' diskette issued quarterly that explains RFB services and updates their catalog.

Computerized Books for the Blind (CBFB) merged with RFB in 1991 and E-Text floppy diskettes are available. The cost will vary depending on the number of diskettes required for the book. These diskettes can be used on IBM and IBM compatibles, Apple and Macintosh computers for large print screen viewing, for Braille output and with a Speech Synthesizer. For further information on titles available call the RFB reference library, (609) 452-0606.

* San Francisco Lighthouse for the Blind, Inc., 745 Buchanan Street, San Francisco, CA 94102.

* Sargent, Welch Corporation, P.O. Box 1026, Skokie, IL 60076-8036. (800) 727-4368, (708) 677-0600. Sargent, Welch Corporation is a dealer handling many instruments with Binary Code Decimal (BCD) digital output. Some of these instruments are spectrophotometers, pH meters, Mettler balances, Ohaus balances, and electronic counters. Instruments such as Port-O-Gram Electronic Ohaus Balance can be made 'talking' with a voice box supplied by American Foundation for the Blind.

* Science Products, Box 888, Southeastern, PA 19399. (800) 888-7400. Manufactures special products such as calculators and various meters with talking output for the visually impaired and has a variety of magnifiers. Staff experience in the fields of digital voice technology and equipment adaptation is extensive.

Sensory Access Foundations, 385 Sherman Avenue, Suite 2, Palo Alto, CA 94306.

* Sensory Interface Systems, Larry Waldron, 4441 Kasson RoadSyracuse, NY 13215. (315) 469-7182.

* Smith-Kettlewell Eye Research Institute, Science Products, Mr. William Gerrey, 2232 Webster Street, San Francisco, CA 94115. (415) 561-1677.. Provides service on modification of equipment to aid visually disabled in the laboratory.

Society for the Blind (in most large cities).

Trace Research and Development Center for Communication, Control and Computer Access for Handicapped Individuals, S-151 Waisman Center, 1500 Highland Avenue, Madison, WI 53705, (608) 262-6966 (voice), (608) 263-5408 (TDD), (608) 262-8848 (Fax).

* Volunteers for the Visually Handicapped (VVH). 8720 Georgia Avenue, Suite 210, Silver Spring, MD 20910. (301) 589-0894. This volunteer group develops procedures for making tapes that a print-handicapped person can use to locate specific information quickly within a recorded text or reference book. There is *Voice Indexing Manual* (1989) on how to learn to use this technique available from VVH as well as a list of titles of books-on-tape.

* Xerox Imaging Systems, Inc., 9 Centennial Drive, Peabody, MA 01960. (800) 421-7323, (508) 977-2000, Fax: (508) 977-2148. Provider of the Kurzweil reading machine.

For the Deaf and Hearing Impaired

Captioned Films for the Deaf Distribution Center, Modern Talking Picture Service, 5000 Park Street, North, St. Petersburg, FL 33709. (800) 237-6213 (V/TDD). The Center produces and distributes captioned films. Educational and theatrical films are catalogued, the majority being for primary and secondary schools on nonscientific subjects. There are a few films which have a sound track synchronized with captions, some of which deal with basic chemistry concepts.

National Technical Institute for the Deaf (NTID). Rochester Institute of Technology, One Lomb Memorial Drive, P.O. Box 9887, Rochester, NY 14623-0887. (716) 475-6400 (V/TDD). Persons who take notes for students with hearing or mobility impairments find that a special non-carbon notepad produced by NTID allows them to make multiple copies at one writing. The notepads and other educational aids are listed in a free catalogue, *New Trends for Instructing Deaf People*, available from NTID Information Office. NTID can also be contacted for interpreters.

A SourceBook Module

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

LIBRARY RESOURCES

Library Resources

The following list is a suggestion of useful classroom resources. You will probably find some of them in your school library or in the rooms of experienced teachers. If you are interested in purchasing your own copy, write or call the publisher for availability and current prices. The asterisks mark those we would recommend for earliest addition to your collection.

- Demonstration books are how-to books for illustrating concepts.
- Content resources give you a better understanding of chemistry, and will be useful to advanced students who want to know more.
- Reference books provide facts you may need about elements, compounds and formulas.
- Pedagogy books provide a perspective of chemistry teachers' role in the classroom and society.
- History selections will add personality to the subject matter.
- Periodicals will help you keep current and answer questions about the newspaper headlines.

I. Books on Demonstrations

*Alyea, H. N. and Dutton, F. B. Tested demonstrations in chemistry (6th ed.). *Journal of Chemical Education,* Division of Chemical Education of the American Chemical Society, Flinn Scientific, Inc., P.O. Box 219, Batavia, IL 60510-0219; (708) 879-6900.

Alyea, Hubert. TOPS (3rd ed.). *Journal of Chemical Education,* Division of Chemical Education of the American Chemical Society, 1991 Northampton St., Easton, PA 18042; (215) 250-7264.

Alyea, Hubert. (1986). *Armchair chemistry: A Programmed Laboratory Manual,* (2nd ed.). Flinn Scientific, Inc., P.O. Box 219, Batavia, IL 60510-0219; (708) 879-6900.

Borgford, Christie L., and Summerlin, Lee R. (1988). Chemical activities. *American Chemical Society,* 1155 Sixteenth St. NW, Washington, DC 20036; (800) 227-5558. (Note: A teacher edition is available.)

Boys, C. V. (1959). *Soap bubbles.* Dover Publications, 180 Varick St., New York, NY 10014. Toll free: (800) 223-3130; or Flinn Scientific, Inc., P.O. Box 219, Batavia, IL 60510-0219; (708) 879-6900.

Brown, Robert J. *333 Science tricks and experiments.* Tab Books, Division of McGraw-Hill, Inc., P.O. Box 40, Blue Ridge Summit, PA 17294-0850; (800) 822-8138.

Brown, Robert J. (1984). *333 More science tricks and experiments.* Tab Books, Division of McGraw-Hill, Inc., P.O. Box 40, Blue Ridge Summit, PA 17294-0850; (800) 822-8138.

Chen, Philip S. (1974). *Entertaining and educational chemical demonstrations.* Chemical Elements Publishing Co., 529 Mission Drive, Camarillo, CA 93010. [Out of print]

Cobb, Vicki. (1972). *Science experiments you can eat.* J.B. Lippincott, Subsidiary of Walters Kluwer U.S. Corp., 227 E. Washington Sq., Philadelphia, PA 19106-3780; (800) 441-4526.

Cobb, Vicki. (1982). *The secret life of hardware.* J.B. Lippincott, Subsidiary of Walters Kluwer U.S. Corp., 227 E. Washington Sq., Philadelphia, PA 19106-3780; (800) 441-4526.

Cobb, Vicki. (1985). *Chemically active.* J.B. Lippincott, Subsidiary of Walters Kluwer U.S. Corp., 227 E. Washington Sq., Philadelphia, PA 19106-3780; (800) 441-4526.

Cobb, Vicki. (1983). *Gobs of goo.* J.B. Lippincott, Subsidiary of Walters Kluwer U.S. Corp., 227 E. Washington Sq., Philadelphia, PA 19106-3780; (800) 441-4526.

Cobb, Vicki. (1979). *More science experiments you can eat.* New York: J.B. Lippincott, Subsidiary of Walters Kluwer U.S. Corp., 227 E. Washington Sq., Philadelphia, PA 19106-3780; (800) 441-4526.

A collection of chemical demonstrations. (1993). *Chem 13 News.* Waterloo, ON: Department of Chemistry, University of Waterloo, N2L 3G1, CANADA; $3.00.

Conkling, John A. (1985). *Chemistry of pyrotechnics.* Marcel Dekker, Inc. [Out of print]

Dreyfus / Woodrow Wilson Institute Materials. Princeton, NJ: Woodrow Wilson National Fellowship Foundation, P.O. Box 642, 08542.

Gardner, Robert. (1985). *Science around the house.* Julian Messner, Division of Silver Burdett Press, Simon & Schuster, Inc., 190 Sylvan Ave., Englewood Cliffs, NJ 07632; (201) 592-2646.

Herbert, D. (1980). Mr. *Wizard's supermarket science.* Random House, Inc., 201 E. 50th St., 31st Floor, New York, NY 10012; (800) 733-3000.

Herbert, D., and Ruchlis, Hy. (1968). Mr. *Wizard's 400 experiments in science.* Book-Lab, P.O. Box 7316, 500 74th St., North Bergen, NJ 07047-1016; (800) 634-4081. (Revised 1983 by David Goldberg).

Herbert, D. (1959). Mr. *Wizard's experiments for young scientists.* Doubleday & Co., Inc., 666 Fifth Ave., New York, NY 10103; (800) 223-6834.

Humphreys, David. *Demonstrating chemistry.* Hamilton, ON: McMaster University, Department of Chemistry, L8S 4M1, CANADA

Idea Bank Collation. c/o S17 Science, Box 1591, Kingston, ON K7L 5C8, CANADA.

Johnson, Mary. (1981). *Chemistry experiments.* Usborne Publishing, 20 Garrick St., London WC2E 9BJ, England.

Mebane, Robert C. and Rybolt, Thomas R. (1985). *Adventures with atoms and molecules* (Vol. I-III). Enslow Publishers, P.O. Box 777, Hillside, NJ 07205; (908) 964-4116.

Meiners, Harry F. *Physics demonstration experiments, 1 and 2.* New York, NY: Ronald Press Co. John Wiley & Sons, Inc., 605 Third Ave., New York, NY 10158-0012; (212) 850 6000. (See p. 653, Vol. 1 for extensive list of physics demonstration books.)

Mullin, Virginia L. (1968). *Chemistry experiments for children.* Garden City, NY: Dover Publications, Inc.; (800) 223-3130.

National Science Teachers Association. *Polymer chemistry.* National Science Teachers Association. Affiliation of American Association for Advancement of Science, 5110 Roanoke Pl., Suite 101, College Park, MD 20740; (301) 474-9227.

Palder, Edward L. (1987). *Chemistry magic.* Kensington, MD: Woodbine House, 10400 Connecticut Avenue, 20895; (800) 843-7323.

Selinger, Ben. (1989). *Chemistry in the marketplace* (4th ed.). Harcourt Brace Jovanovich, Inc., 1250 Sixth Ave., San Diego, CA 92102. Toll free: (800) 346-8648. Available from Flinn Scientific, Inc., P.O. Box 219, Batavia, IL 60510-0219; (708) 879-6900.

Shakhashiri, Bassam Z. (1983). *Chemical demonstrations, 1.* University of Wisconsin Press, 114 North Murray Street, Madison, WI 53715-1199; (608) 262-8782.

Shakhashiri, Bassam Z. (1985). *Chemical demonstrations, 2.* University of Wisconsin Press, 114 North Murray Street, Madison, WI 53715-1199; (608) 262-8782.

Shakhashiri, Bassam Z. (1989). *Chemical demonstrations, 3.* University of Wisconsin Press, 114 North Murray Street, Madison, WI 53715-1199; (608) 262-8782.

Shakhashiri, Bassam Z. (1992). *Chemical demonstrations, 4.* University of Wisconsin Press, 114 North Murray Street, Madison, WI 53715-1199; (608) 262-8782.

Shamus, H. *Great experiments in physics.* Holt Rinehart & Winston, Inc., Division of Harcourt Brace Jovanovich, Inc., 301 Commerce St., Fort Worth, TX.; (800) 447-9479.

*Summerlin, Lee R. and Ealy, James L. Jr. (1985). *Chemical demonstrations. A sourcebook for teachers.* American Chemical Society, 1155 Sixteenth St., NW, Washington, DC 20036; (800) 227-5558.

*Summerlin, Lee R., Borgford, Christie. L., and Ealy, Julie B. (1988). *Chemical demonstrations (2nd ed.). A sourcebook for teachers, 2.* American Chemical Society, 1155 Sixteenth St., NW, Washington, DC 20036; (800) 227-5558.

Sutton, M. *Demonstration experiments in physics.* McGraw-Hill Publishers, 1221 Avenue of the Americas, New York, NY 10020; (800) 262-4729.

Talesnick, Irwin. (1982). *Demonstrations in chemistry.* Kingston, ON: Queen's University, K7L 3N6, CANADA.

Tocci, Salvatore. (1985). *Chemistry around you.* Prentice Hall Press. A division of Simon & Schuster, Inc., 15 Columbus Circle, New York, NY 10023; (800) 223-2348.

VanCleave, Janice P. (1989). *Chemistry for every kid.* New York, NY: John Wiley & Sons, Inc., 605 Third Ave., 10158-0012; (212) 850-6000.

Vivian, Charles. (1967). *Science experiments & amusements for children.* Garden City, NY: Dover Publications, Inc.; (800) 223-3130.

II. Books for the Professional Library

A. Content Resources

Bailar, J. C. *et al.* (1989). *Chemistry* (3rd ed.). Harcourt Brace Jovanovich, Inc., 1250 Sixth Ave., San Diego, CA 92101; (800) 346-8648.

*Bohinski, R. (1987). *Modern concepts in biochemistry.* Allyn and Bacon, Division of Simon & Schuster, 160 Gould St., Newton Heights, MA 02194; (617) 455-1200.

Brown, T. L, Lemay, H. E. and Bursten, B. (1991). *Chemistry, the central science.* (5th ed.). Prentice Hall, Division of Simon & Schuster, Inc., 15 Columbus Cir., New York, NY 10023; (800) 922-0579.

**ChemCom: Chemistry in the community.* (2nd ed.) (1993). Kendall/Hunt Publishing Company, 2460 Kerper Blvd., Dubuque, IA 52001; (800) 338-5578.

Ebbing, D. (1993). *General chemistry.* (4th ed.) Houghton-Mifflin, One Beacon Street, Boston, MA 02108; (800) 225-3362.

Emsley, J. (1991), *The Elements* (2nd ed.). Clarendon Publishing Co., 267 Walker Ave., Clarendon Hills, IL 60514; (708) 323-0625.

Gardner, M., Heikkinen, H., and Smith, P. (1990) *Laboratory assessment builds success.* Institute for Chemical Education, Lawrence Hall of Science, University of California, Berkeley, CA 94720; (510) 642-1016.

Gillespie, R. *et al.* (1989). *Chemistry* (2nd ed.). Allyn and Bacon, Division of Simon & Schuster, 160 Gould St., Newton Heights, MA 02194; (617) 455-1200.

Goldberg, D. E., *et. al.* (1988). *Schaum's 3000 solved problems in chemistry.* McGraw Hill Book Publishers, 1221 Avenue of the Americas, New York, NY 10020; (800) 262-4729.

*Hart, H. (1991). *Organic chemistry: A short course.* (8th ed.). Houghton Mifflin, One Beacon St., Boston, MA 02108; (800) 225-3362.

Hill, J. W. *Chemistry for changing times* (6th ed.). (1992). Macmillan Publishing Company, 866 Third Avenue, New York, NY 10022; (800) 257-5755.

*Kotz, J. C., and Purcell, K. F. (1991). *Chemistry and chemical reactivity* (2nd ed.). Saunders College Publishing, Division of Harcourt Brace Jovanovich, Inc., 465 Lincoln Dr., Troy, MO 63379; (314) 528-8131.

McQuarrie, D. A., and Rock, P. A. (1991). *General chemistry* (3rd ed.). W.H. Freeman and Company, 41 Madison Ave., New York, NY 10010; (212) 576-9400.

Petrucci, R. H., Harwood, W. S. (1993) *General chemistry.* (6th ed.). Macmillan Publishing Co., Inc., 866 Third Ave., New York, NY 10022; (800) 257-5755.

*Pimentel, G. & Coonrod, J, (1987) *Opportunities in chemistry today and tomorrow.* National Academy Press, 2101 Constitution Ave., NW, Washington, DC 20418; (800) 624-6242.

Ragsdale, R. (1987), *Problems in chemistry: questions and solutions.* Saunders College Publishing, Division of Harcourt Brace Jovanovich, Inc., 465 Lincoln Dr., Troy, MO 63379; (314) 528-8131.

Skoog, D., Holler, F., and West, D. (1990). *Analytical chemistry: An introduction* (5th ed.). Saunders College Publishing, Division of Harcourt Brace Jovanovich, Inc., 465 Lincoln Dr., Troy, MO 63379; (314) 528-8131.

Weinberg, S. (1977). *The first three minutes.* Basic Books, Inc., Division of HarperCollins Pubs., Inc., 10 E 53rd St., New York, NY 10022; (800) 242-7737.

B. Reference Books

Budavari, S. (Ed.). (1989). *The Merck index (*11th ed.). Merck & Co, Inc., Rahway, NJ 07065; (201) 855-4558.

*CRC *handbook of chemistry and physics.* CRC Press, Inc., 2000 Corporate Blvd., NW, Boca Raton, FL 33431; (800) 272-7737.

Dean, J. A. (Ed.). (1985). *Lange's handbook of chemistry.* McGraw-Hill Publishers, 1221 Avenue of the Americas, New York, NY 10020; (800) 262-4729.

C. Pedagogy and Philosophy

Amondson, K. J. (1991). *Teaching values and ethics.* Arlington, VA: American Association of School Administrators.

Bogner, Donna. (1989). *Starting at ground zero: A methods book for teaching first-year chemistry*.Hutchinson, KS: Genie Publications, 4318 South Willison Road, 67501.

Collette, A. T. and Chiapetta, E. L., Eds. (1989). *Science instruction in the middle and secondary schools* (2nd ed.). Columbus, OH: Merrill Publishing Co.

Educating Americans for the 21st century. (1983). National Science Board, Commission on Precollege Education in Mathematics, Science, and Technology, National Science Foundation, Washington, DC 20550.

Flavell, J., Miller P. and Miller, S. (1993). *Cognitive development* (3rd ed.). Englewood Cliffs, NJ: Prentice Hall.

Grubbs, W., Davis, G., Lum, J., Pihal, J. and Morgraine, C. (1991). *The cunning hand, the cultured mind: Models for integrating vocational and academic education*. Berkeley, CA: National Center for Research in Vocational Education.

Hassard, J. (1992). *Minds on science: Middle and secondary school methods*. New York: Harper Collins Publishers.

Hazen, R. and Trefil, J. (1991). *Science matters*. Doubleday & Co., Inc., 666 Fifth Ave., New York, NY 10103; (800) 223-6834.

Hull, D. and Poinell, D. (1991). *Tech Prep associate degree: A win / win experience*. Waco, TX: Center for Occupational Research and Development.

Kuhn, T. (1974). *The structure of scientific revolutions*. University of Chicago Press, 5801 Ellis Ave., Chicago, IL 60637; (800) 621-2736.

Lawson, A., Abraham, M., and Renner, J. (1989). *Monograph 1. A theory of instruction: Using the learning cycle to teach science concepts and thinking skills*. National Association for Research in Science Teaching, Dr. John Staver, Executive Secretary, Bluemont Hall, Kansas State University, Manhattan, KS 66506-5334.

Lickona, T. (1991). *Educating for character: How our schools can teach respect and responsibility*. New York, NY: Bantam Books.

Medawar, P. (1984). *The limits of science*. New York: Harper & Row. [Out of Print]

*Morrison, P. (1989). *The ring of truth: an inquiry into how we know what we know*. Random House, Inc., 201 E. 50th St., 31st Floor, New York, NY 10022; (800) 733-3000.

Nicholls, J. (1989). *The competitive ethos and democratic education*. Cambridge, MA: Harvard University Press.

Resnick, L. B. (Ed.) (1989). *Knowing, learning, and instruction: Essays in honor of Robert Glaser*. Lawrence Erlbaum Associates, Inc., 365 Broadway, Hillsdale, NJ 07642.

**Science for all Americans.* (1989). Washington, D.C.: American Association for the Advancement of Science, 1333 H St., NW, 11th Floor, Washington DC 20005; (202) 326-6446.

Sharon, S. (Ed.) (1992). *Cooperative learning: Theory and research.* New York, NY: Praeger.

Tobias, S. (1990). *They're not dumb, they're different: Stalking the second tier.* Research Corporation, 6840 East Broadway Blvd., Tucson, AZ 86710-2815.

Trowbridge, L. W. and Bybee, R. W. (Eds.) (1990). *Becoming a secondary school science teacher* (5th ed.). Columbus, OH: Merrill Publishing Co.

D. History of Science

Asimov, Isaac. (1982). *Asimov's biographical encyclopedia of science and technology.* Doubleday & Co., Inc., 666 Fifth Ave., New York, NY 10103; (800) 223-6834.

Gamow, George. (1966). *Thirty years that shook physics.* Doubleday & Co., Inc., 666 Fifth Ave., New York, NY 10103; (800) 223-6834.

Hoffmann, Roald and Torrence, Vivian. (1993). *Chemistry imagined: Reflections on science.* Washington, DC: Smithsonian Institution Press.

Ihde, Aaron J. (1984). *The development of modern chemistry.* Garden City, NY: Dover Publications.

Jaffe, B. (1948). *Crucibles: The story of chemistry.* Simon & Schuster, 1230 Avenue of the Americas, New York, NY 10020; (212) 698-7000.

James, Laylin K. (Ed.) (1993). *Nobel laureates in chemistry: 1901–1992.* Washington, DC: American Chemical Society, History of Modern Chemical Sciences Series. $34.95 (paper).

Leicester, Henry M. (1956). *The historical background of chemistry.* Garden City, NY: Dover Publishing Co.

Roberts, R. (1989). *Serendipity: Accidental discoveries in science.* John Wiley & Sons, Inc., 605 Third Ave., New York, NY 10158-0012; (212) 850-6000.

*Salzberg, H.W. (1991). *From caveman to chemist: Circumstances and achievement.* American Chemical Society, 1155 Sixteenth St., NW, Washington, DC 20036; (800) 227-5558.

Travis, Anthony S. (1993). *The rainbow makers: The origins of the synthetic dyestuffs industry in western Europe.* Bethlehem, PA: Lehigh University Press.

Watson, J. (1969). *The double helix.* New American Library, 375 Hudson St., New York, NY 10014-3657; (212) 366-2000.

III. Periodicals

**ChemMatters.* Office of High School Chemistry, American Chemical Society, 1155 Sixteenth St., NW, Washington, DC 20036; (800) 227-5558. $5.50 per year.

Chemunity News. Office of High School Chemistry, American Chemical Society, 1155 Sixteenth St., NW, Washington, DC 20036; (800) 227-5558. Free.

Chem 13 News. Department of Chemistry. University of Waterloo, Waterloo, Ontario. N2L 3G1, Canada; (519) 885-1211, extension 2505. $14.00 per year.

The Journal of Chemical Education. Journal of Chemical Education, Subscription and Book Order Department, 1991 Northampton St., Easton, PA 18042; (215) 250-7264. $34.00 per year.

WonderScience. Office of Pre-High School Chemistry. American Chemical Society, 1155 Sixteenth St., NW, Washington, DC 20036; (800) 227-5558. $4.00 per year.

IV. Media Resources

A. Videotapes and Videodiscs

a. World of Chemistry Videotapes

The High School edition is available from *WINGS for Learning / SUNBURST,* 101 Castleton Street, Pleasantville, NY 10570; (800) 321-7511; (914) 747-3310; (914) 747-4109 (FAX). Print materials accompanying each video segment offer ideas, demonstrations, laboratory experiments, student viewing guide, and student research projects. There are ten 30-minute videotapes, each containing two programs.

The College edition is available from Corporation for Public Broadcasting, The Annenberg/CPB Project, c/o *Intellimation,* P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) LEARNER. There are 26 half-hour programs. Print materials available are a textbook, a student study guide, a laboratory manual, and an instructors manual.

High School Videotapes

1. Color/Measurement
2. States of Matter/The Mole
3. The Atom/The Periodic Table
4. Chemical Bonds/Organic Chemistry
5. Molecular Fingerprints/Water

6. The Atmosphere/The Driving Forces
7. Chemical Kinetics/Catalysts
8. Acids and Bases/Chemistry of the Earth
9. Oxidation and Reduction/Metals
10. Polymers/Biochemistry

b. MECC, 3490 Lexington Avenue North, St. Paul, MN 55126-8097.

c. *The Videodisc Compendium*, Future Systems, Inc., P.O. Box 26, Falls Church, VA 22046; (703) 241-1799.

d. *Recycle This!* Dow Plastics and Dow Brands. (800) 441-4369. Free videotape on recycling.

e. ERIC, Clearinghouse for Science, Mathematics and Environmental Education. The Ohio State University, College of Education, 1200 Chambers Road, 3rd floor, Columbus, OH 43212.

f. *Doing Chemistry* Videodiscs, Education Division, American Chemical Society, 1155-16th Street, NW, Washington, DC 20036; (800) 227-5558.

B. Computer and Multimedia Courseware

Conduit, University of Iowa, Oakdale Campus, Iowa City, IA 52242; (800) 365-9774.

Coronet/NTI Film and Video, 108 Wilmot Road, Deerfield, IL 60015; (800) 621-2131; (708) 940-3640 (FAX).

Falcon Software, P.O. Box 200, Wentworth, NH 03282; (603) 764-5788. Source of such useful software as *Chem$_1$ Problem Proctor* (Stephen K. Lower), *Introduction to General Chemistry* (Smith, S., Chabay, R., and Kean, E.) and *Exploring Chemistry IV, CD-ROM Interactive Video Laboratory* (Smith, S., Jones, L. and Gammon, S.).

Films for the Humanities and Sciences, P.O. Box 2053, Princeton, NJ 08543; (800) 257-5126.

JCE: Software. Department of Chemistry, University of Wisconsin/Madison, Madison, WI 53706-1396; (608) 262-5153; (608) 262-0381 (FAX).

Project SERAPHIM. Department of Chemistry, University of Wisconsin/Madison, Madison, WI 53706-1396; (608) 263-2837; (608) 262-0381 (FAX). Catalog is $10.00; Membership, which includes the catalog and periodic updates, is $20.00 per year.

QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.

SYNAPS Company, 334 S. Cotner Blvd., Lincoln, NE 68510; (402) 489-0667. CD-ROMs and Videodiscs; custom programming and instruction.

TV Ontario Video, 143 West Franklin Street, Suite 206, Chapel Hill, NC 27516; (919) 380-0747.

Ward's Natural Science Establishment, Inc., 5100 West Henrietta Road, P.O. Box 92912, Rochester, NY 14692-9012; (716) 359-2502; (800) 962-2660; (716) 334-6174 (FAX). Distributors of the CHEM Study Films and Videos.

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VII. Other Resources

Chemical Heritage Foundation, 3401 Walnut Street, Philadelphia, PA 19104-6228; (215) 898-4896; (215) 898-3327 (FAX). *Chemical Heritage* is a free newsletter published by this foundation. It contains historical vignettes about chemistry, book reviews and bibliographical information on the history of chemistry.

DivCHED Newsletter, published three times per year by the Division of Chemical Education, American Chemical Society. Contains news and announcements pertinent to chemical education nationwide; abstracts of Division papers that are presented at each national ACS meeting. Subscription is part of DivCHED membership: $10.00 per year. Mary Virginia Orna, Treasurer, College of New Rochelle, New Rochelle, NY 10805; (914) 654-5302; (914) 654-5387 (FAX).

The Flinn Catalog, Flinn Scientific, Inc., P.O. Box 219, Batavia, IL 60510-0219; (708) 879-6900. A universal resource for the high school teacher. It contains safety information and many other Flinn "goodies." Be sure to get on their mailing list.

Other resources, references and media are incorporated into each *SourceBook* module.

The *Library Resources* module was compiled by Mary Jo Grabner, C. Leonard Himes and Mary Virginia Orna. Material was supplemented by M. Lynn James and the *GuideBook* writing team.

PROFESSIONAL ORGANIZATIONS

A SourceBook Module

Version 2.0 1994

Funded in part under
National Science Foundation
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ChemSource Project Principal Investigator:
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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Professional Organizations

The following organizations welcome teachers as members. Their publications will be very useful in keeping current on chemistry content and informing you of such things as summer institutes and workshops. They are also sources of grants for project development. Attending their meetings will put you in touch with other professionals in your area and nationally. Membership in at least one of these organizations should be made a high priority.

In addition to national organizations, there are also state and local organizations and even users' groups for specific curricula. Ask your textbook publisher, department chair, school district or state/local science supervisor to learn more. Such groups are likely to sponsor useful workshops in your area and provide a network for professional development and assistance in dealing with chemistry teaching problems.

AlChemE, Department of Chemistry, SUNY/Purchase, Purchase, NY 10577. (914) 251-6691.

A chemistry teachers' enrichment group for lower New York State teachers.

American Chemical Society, 1155-16th Street, NW, Washington, DC 20036. (202) 872-4382. (Martha Turches, High School Office)

Yearly Membership: $99.00. Chemistry professional organization; dues includes subscription to weekly magazine, *Chemical and Engineering News*.

American Association of Physics Teachers (AAPT), 5112 Berwyn Rd., College Park, MD 20740.

Chemical Heritage Foundation (formerly CHOC). 3401 Walnut Street, Philadelphia, PA 19104-6228. (215) 898-3327.

Free Newsletter; write to be put on their mailing list.

Chemistry Teachers' Club of New York, (Chem Club), 207 Lincoln Place, Eastchester, NY 10707-2005. (914) 961-8882.

Yearly Membership: $10.00. Monthly newsletter; meeting activities; monthly meetings in New York City.

ChemNexions, C/o Bannon, 303 Fifth Avenue, Saint James, NY 11780. (516) 862-9405.

A chemistry teachers' group on Long Island.

Division of Chemical Education, ACS, 16 Hemlock Place, New Rochelle, NY 10805. (914) 654-5302.

Yearly Membership: $10.00. Newletter published three times yearly; information on conferences, meetings, all the latest in chemical education.

Institute for Chemical Education, Dept. of Chemistry, University of Wisconsin–Madison, 1101 University Ave., Madison, WI 53706. (608) 262-3033.

National Association for Research in Science Teaching (NARST), 219 Bluemont Hall, Kansas State University, Manhattan, KS 66506.

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National Science Teachers Association, 1742 Connecticut Ave., NW, Washington, DC 20009. (202) 328-5800.

Yearly Membership: $35.00. Three area meetings and one national meeting annually; publishers of *The Science Teacher* and *Journal of College Science Teaching*, among others.

New England Association of Chemistry Teachers, C/o C. Francis Egan, The Hotchkiss School, Lakeville, CT 06039.

Yearly Membership: $10.00. Annual conference and many meetings during the year; publishes *NEACT Journal*; membership not limited to residents of New England.

Science Teachers Assn. of New York State (STANYS), 491 Oakdale Road, East Meadow, NY 11554. (516) 783-5432.

Yearly Membership: $29.00. Annual conference held in Ellenville, NY; newsletter.

A SourceBook Module

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ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

SMALL-SCALE CHEMISTRY

Small-Scale Chemistry

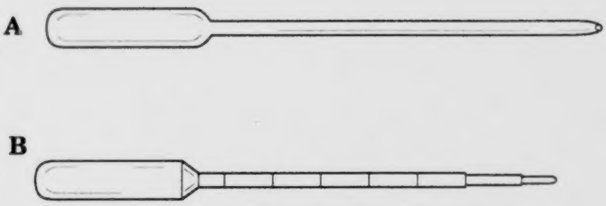
INTRODUCTION

Recent years have shown an increased sensitivity to the hazards of chemicals when used by teenagers in high school laboratories. Combined with limited space, administrative concerns, and mandates of costly waster disposal, many teachers have opted to incorporate scaled-down chemistry activities in the curriculum. Scaling down takes the form of microscale or small scale. Microscale (microchemistry) refers to chemistry performed on microgram and submicrogram quantities. Small-scale chemistry employs a modern design philosophy and a small, inexpensive, plastic apparatus to create a learning laboratory that fosters creativity, invention, and problem solving.

Proponents of both scales stress "less is more." Users find less need for sophisticated apparatus and large quantities of chemicals, less space needed for laboratories or storage, less chemical waste generated, less laboratory clean up time, and less student learning anxiety. Moreover, these users realize more student attention; more thinking and learning among students; better student skills development; enhanced safety; the ability to complete more laboratory activities; and additional time to finish, interpret, and/or repeat laboratories.

The two most important items (apparatus) used in the laboratory are the Beral pipet and the reaction plate (Figure 1).

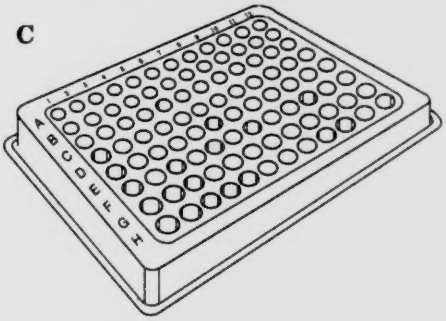
Figure 1. (A) Thin stem pipet, (B) graduated pipet, and (C) reaction plate.



The Beral pipet, made of polyethylene, can have a thin stem, a microtip, an extra large bulb, or it can be graduated. A Beral pipet can be filled with 4-5 mL solution and can deliver 2-50 drops per milliliter, depending on the stem size. (The pipet stem can be pulled and stretched to change the volume of the drop delivered.) The graduated pipet will deliver volumes of 1 mL or less, enabling it to be used as a microburet. Pipets can be labeled and stored vertically (tip up) in a small space for use the following year.

A unique use of the Beral pipet with the extra large bulb is as a microchromatography column. Once the bulb is cut in half, a tiny cotton plug can be inserted into the pipet to retain the medium. After the pipet is filled with the medium, the bulb acts as a funnel and reservoir for the eluting solvent.

The reaction plate is really a set of micro testtubes. The plates are made of polystyrene and are flat, with each well having a slight rim. The well capacity varies from 0.3 mL to 9.6 mL. Reaction plates can be obtained in a variety of sizes, including 6, 24, 48, and 96 wells.



Sample identification is easy because the plate rows are marked by letters and numbers. Plates are easy to clean, and stubborn residues can be removed by a cotton swab, rinsed with water, and shaken dry. The plates can be stored by stacking.

Culture flasks and Pasteur pipets (Figure 2) can be used as needed. Culture flasks are made of polystyrene and can be used as containers for large quantity solution storage. These flasks are constructed so that they stack on their side, allowing many containers to be stored in a limited space. Pasteur pipets are disposable Pyrex glass. The nondelivery end will accept a soda straw that, when bent, will provide a way of controlling the delivery of solutions. The narrow tip length can also be removed (a length can be broken off), the small end of the pipet plugged with glass wool, and the pipet used as a microfilter.

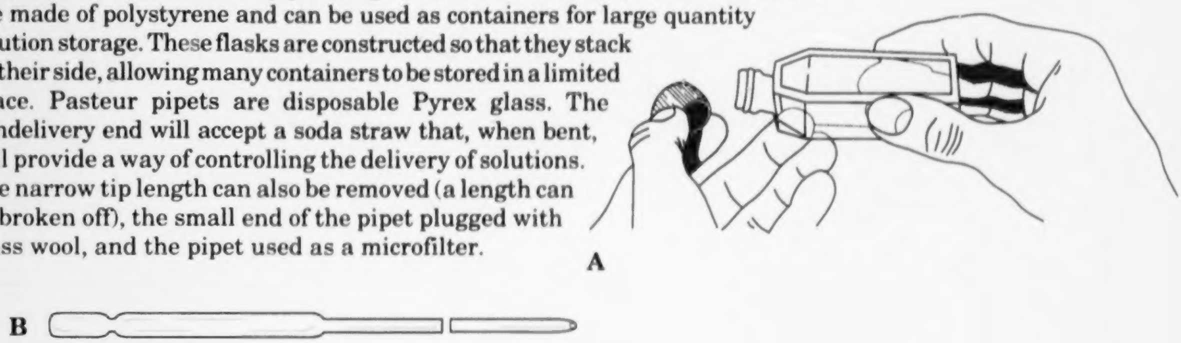


Figure 2. (A) culture flask and (B) Pasteur pipet.

Many *SourceBook* laboratory activities are small-scale versions, but some are microscale or can be converted to microscale. Many commercial laboratory texts are available with a variety of microscale experiments (these are listed below). Materials are available from several chemical suppliers (Flinn Scientific, Micro Mole Scientific, *etc.*). As teachers become involved in scaling down, new laboratory activities are certain to appear. You can become more proficient and comfortable with scaled-down chemistry by attending workshops provided by the Institute for Chemical Education, national and regional meetings of the National Science Teachers' Association, your state Science Teachers' meeting, or at biennial meetings of the American Chemical Society's Division of Chemical Education. Several journals (*Journal of Chemical Education*, *Journal of College Science Teaching*, *etc.*) often have scaled-down chemistry activities for use in the high school laboratory. The authors of the texts mentioned below are reputable experts in microchemistry and could be contacted for assistance.

SourceBook editors have reviewed the publications below for microchemistry and small-scale activities. We have grouped these and keyed them to the laboratory activities in *SourceBook*. For example, Activity 1 in the ACID module is "Classifying Substances Based on Their Reactions." Reference (1) on page 12 contains a small-scale activity on "Natural Indicators." In some cases, we did not find a suitable small-scale chemistry activity and have used the notation *** to so indicate. Perhaps you would like to develop a small-scale or microscale activity and submit it to the editors to consider for including in *SourceBook*.

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Volume 1

ACID	Reference/Page
Activity 1: Classifying Substances Based on Their Reactions	
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The Rainbow Lab	1/9
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Is Your Vinegar Correctly Labeled?	4/115
Changing Partners; Cation Exchange Chromatography	4/123
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Analyzing Commercial Vinegar (Chapter 20)	5 /45
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Section B: Indicator Color Probes for Acids and Bases	6/150
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Instructional Resources for Preservice and Inservice Chemistry Teachers

CHEMICAL SAFETY

Chemical Safety

INTRODUCTION

Much of the information found in this document is based on, taken from, or otherwise follows the format established in the publication *Safety in Academic Chemistry Laboratories,* 5th ed. (1990). (SACL), *(1)* published by the American Chemical Society. When sections are excerpted directly, they will be in paragraphs offset from the text, so that readers will know which sections are general (those from that publication), and which are written specifically for the current audience. The guidelines here have been prepared by the ACS Joint Board/Council Committee on Chemical Safety (BCCCS) in cooperation with the governing committee of the *ChemSource* Project. These guidelines can be adapted to practices in all school laboratories using chemicals, including student research areas, storage areas for chemicals and equipment, and the classroom and laboratory.

It may be true that "laws of the land" demand an effective safety program, as indicated in SACL, but concern for education, health and safety should be a fundamental part of any academic program even without laws. The law is helpful in assuring that the classroom teacher or laboratory helper has guidelines and regulations on which to base judgment with regard to laboratory practice. The practice of chemistry and the learning of chemistry are rooted in laboratory experience. That experience should not be denied to students because of the threat of a law. Instead, regulations should be used as tools to help students practice chemistry safely and really become educated. Implied in that step is that students must learn how to judge the potential hazards in a particular operation or activity, and to devise means by which potential risk may be reduced to an acceptable level.

An effective safety program requires strong support from the top administrative officer and will actively involve students, staff, faculty, and administrators. Without the enthusiastic support of the top administrative officer, who is ultimately responsible for safety within any facility, an effective safety program is rarely achieved. A safety officer, appropriately trained and qualified, is essential. However, even the best safety officer cannot relieve the head of the laboratory, the faculty and the students of their responsibility for the safety of operations under their jurisdiction. A good safety program includes constant training and reminders of the common hazards to which laboratory workers may be exposed, as well as attention to the materials, equipment and facilities with which they work.

> *Good practice requires mandatory safety rules and programs. To achieve safe conditions for students, faculty, and visitors, the program must include regular safety inspections at intervals of no more than three months (and shorter periods for certain types of equipment), disposal procedures to ensure that waste chemicals are properly handled, and regular safety programs to ensure that sufficient full-time personnel are trained in the proper use of emergency equipment and procedures, and regular monitoring of the performance of ventilation systems (2).*

This document is directed toward helping teachers, laboratory workers, staff, and school personnel develop and establish strong and useful safety habits.

Teachers, often under-prepared, are asked to provide more challenging courses, while at the same time they are being warned of liability consequences if known hazardous chemicals are used improperly.

Wahl *(2)* comments:

> *Regulations covering the use and disposal of chemicals are proliferating, but teachers are given little guidance and less time and money to properly implement them. Facilities for high school laboratories and chemical storerooms are generally inadequate and rarely secure.*

It is the intent of this document, and Wahl's, to help provide some of the information necessary to provide a strong, stimulating, and safe laboratory experience for students. We want students to feel that they had a good experience in the laboratory, and we especially want them to know that they learned something in the laboratory setting. We must not deny the laboratory experience to our students.

RESPONSIBILITY FOR SAFETY

Introduction

The information contained in this document should not be considered definitive. We believe that if teachers know something about interpretations of the law, and its application to other cases, then they may be forewarned and forearmed.

Responsibilities

Every individual in the chemistry laboratory (students, visitors, workers, staff, faculty) and appropriate individuals in the governance hierarchy of the school and school district must participate in management of the safety program in order to assure as safe an environment as possible. The faculty and staff of the chemistry or science department are responsible for the administration of the safety program, but it can be successful only in so far as the administrative support is strong.

Untrained students or visitors should not be permitted to work with chemicals inappropriate to their level of proficiency in the school laboratory. A laboratory instructor should be responsible for no more than twenty-four students at one time in order to provide support for the simplest and safest laboratory programs. The instructor should be in the laboratory for the entire laboratory period. [One reviewer of this material has suggested that a teaching assistant, if possible, or other responsible adult be designated as a "deputy" in case of an emergency where the teacher must leave the room. This seems to be a common-sense action, but one that is untested as to any legal standing.] Some teachers have instituted programs by which seniors are trained as assistants and asked to serve in that capacity with other level students. One such program has been described in *ChemUnity (3)*, and safety concerns that resulted have been fully reported. The best circumstance, of course, is to have the teacher present, but if that is not possible, this may be a suitable **short-term** (a few minutes at most) solution.

A school district, principal or other responsible authority should not require teachers to work in an environment (with too many students) that cannot be well managed, particularly with regard to safety issues. It is also true that the school district must provide the tools by which the teacher can maintain a safe laboratory. [Teachers should use this document as the basis for developing a safety program to provide any equipment deemed necessary for safety, but for which the school district has not planned.] This document may be the basis for furthering the cause of safety in the school laboratory.

According to SACL, every instructor or laboratory supervisor must:

1. *Set a good example by:*
    - *observing all rules and recommendations,*
    - *wearing suitable protective equipment and clothing,*
    - *being enthusiastic about the establishment of a safe environment.*
2. *Be alert for unsafe conditions.*
3. *Conduct frequent and thorough inspections of the laboratory and affiliated spaces.*
4. *Maintain discipline and enforce rules with regard to chemical safety.*
5. *Take effective corrective action promptly when an unsafe condition is determined to exist.*

[**Note**: In cases where corrective action requires budgetary or other administrative support, or when safe class size is exceeded, the instructor is obligated to inform the appropriate persons, **in writing**, concerning the need, and to continue to bring the need to the attention of the appropriate persons until it is corrected. To do anything less is to be unfair to oneself, let alone to students in the laboratory. Keep copies of this correspondence in a file away from the school.]

6. *Assume responsibility for visitors and require that they follow the same rules as other laboratory workers and are escorted or supervised at all times.*

7. *Review all procedures for possible health, safety, and environmental problems before the work is begun.*

8. *Maintain a readily available file of publications on chemical safety and encourage its use.*

9. *Obtain or generate a* ***material safety data sheet*** *(MSDS) for each chemical used in the laboratory, and understand the appropriate information found in that document.*

10. *Until the MSDS sheets become standardized and the information they contain has been checked for accuracy, the best source for information on chemicals and their hazards includes The* Merck Index. *This, or other books like it, should be readily available and consulted often.*

Not only is there pressure to provide adequate safety precautions and training in the laboratory, but there is increasing recognition of the dangers of prolonged, low-level exposure to chemicals. While students probably will not be affected by such exposure since they do not spend much time in the laboratory, it is a potential problem of which the instructor or staff member should be aware.

Scope

The law in which nearly all safety regulations are grounded is the act establishing the Occupational Safety and Health Administration (OSHA), the National Institute of Occupational Safety and Health (NIOSH), and the description of their responsibilities. Those laws and regulations developed by and enforced by those and other agencies, are written basically as if describing relationships between an employer and an employee. Students, of course, do not fit the exact definitions, but their rights are no less important and the laboratory director or instructor must insure that their interests are protected.

Disclaimers

It is good practice to always include some form of disclaimer in documents prepared. For instance, if directions are distributed for a new experiment, demonstration, or student activity, those directions should contain information concerning the safety practices appropriate to that procedure, and should disclaim responsibility if the procedures lead to unsuspected consequences. Keep in mind that activities which are not deemed safe for students should be altered in their delivery mode, have additional safety factors added, or be deleted entirely.

A sample disclaimer is shown here.

> **DISCLAIMER**
>
> **No chemical procedure can be guaranteed "safe." All experiments recommended here appear to present minimum hazards. However, there is no guarantee, expressed or implied, that an experiment or procedure will cause no injury. Each teacher selecting an experiment should try it out carefully under controlled conditions before using it as a class activity. Any necessary safety precautions should be added to the material given to the students.**

The American Chemical Society, the ACS Joint Board/Council Committee on Chemical Safety, the Steering Committee for the *ChemSource* project, and others are interested in promoting safety in the high school laboratory. This document should not be considered either complete or binding. The persons and organizations connected with production of this information are not responsible for its use or misuse.

Liabilities

As indicated in the earlier section, a teacher is responsible for:

A. instruction
B. supervision
C. safety
D. maintenance

Foreseeability refers to the expectation that the well-trained professional should be expected to "foresee" dangerous conditions. This requires that the teacher carefully consider each step in any demonstration or laboratory procedure and consider the consequences of student behaviors that do not conform to the described process—for instance, the instructor should play a "what if" game (*i.e.,* "What if the student poured solution A into solution B at this point?").

The argument has been made that students, by definition, are inexperienced. Therefore, the teacher must guard them against possible hazardous consequences of their behavior. While the philosophical implications cannot be explored fully, students most surely are inexperienced; the entire teaching process is one in which we are trying to provide suitable experiences for learning to take place. Yet no activity is absolutely "safe," or without potential hazardous consequences. We need to foresee those possible consequences and reduce the potential for them as much as possible.

Also implied in this notion is that printed instructions **cannot** take the place of verbal instruction, since printed instructions do not allow for questioning nor a recognition on the part of the teacher that the words may not be fully understood.

Negligence has two references in the high school laboratory. The first is reflected in the supervisor's (principal, department chair, *etc.*) obligations to the staff member, teacher, instructor, *etc.* (who is the employee). The second is in reference to the relationship between the instructor and the student.

In the case of the instructor/student relationship, the instructor may be negligent if he or she takes illegal actions that result in a hazardous situation, or if the instructor does not exercise reasonable care in managing the laboratory.

The **Prudent Person Theory** refers to the notion that any activities that are unsafe, and for which the results could have been predicted by a well-trained instructor constitute the standards of a prudent person, and establish the "baseline" for safety practice against which an action will be judged. That is, the answer to the question, "Could a prudent person, trained with appropriate background, have predicted this consequence?," marks the demarcation line between negligence and prudent behavior on the part of the instructor.

Standard of care is a notion that also impacts on considerations of appropriate safety standards. The standard of care refers to what an ordinary, prudent and reasonable instructor would do under the same conditions.

SAFETY RULES FOR THE CHEMICAL TEACHING LABORATORY

Introduction

Many of the rules for the laboratory in the elementary, middle school, junior high, or high school are "common sense" rules. Many are specified as part of federal, state, local, school district, or other agency guidelines. Many are simply an application of good classroom or laboratory management techniques.

Under the OSHA regulations, a specific set of safety rules must be developed and communicated clearly to laboratory workers and students. Note here that the regulation is written to indicate that the school district must meet these guidelines for each worker, including the teacher who is in the laboratory situation. On the other hand, each teacher has an obligation to provide appropriate safety information to each student and the class as a whole, whenever appropriate. This should be done not only in a general way at the start of the school year, but also for each experiment prior to its being completed in the laboratory.

Safety meetings must be held at regular intervals with all faculty and staff. The school should have a safety program, part of which is specific to the chemistry laboratory. A safety committee should be established.

The use of a safety presentation and assignment for each student is strongly recommended. Each student should attest by signature that he or she has read and understands the safety rules. The instructor should keep those signed contracts on file (somewhere other than the laboratory). While this does not guarantee that the teacher conducts a safe laboratory, and does not exonerate the teacher from potential action, it will help the student to understand that the teacher takes safety seriously, and the contracts can be used as reminders for students when appropriate. The existence of those signed contracts can be used to argue that the teacher "intended" to conduct a safe laboratory, and made an effort to do so. The rules must be rigorously and impartially enforced. Willful noncompliance should result in dismissal or suspension from the laboratory.

The following are suggested as rules for persons in the laboratory:

1. Always wear proper eye protection in chemical work, handling, and chemical storage areas. All goggles or face shields should meet the ANSI (American National Standards Institute) Code Z87.1. *(4)*. Contact lenses should normally not be worn in the laboratory. Special precautions must be taken if they are worn.
2. Always know the physical and chemical properties of the materials used in an experiment or demonstration (*e.g.*, corrosiveness, flammability, reactivity, toxicity, and stability).
3. Always wear appropriate protective clothing, including shoes with enclosed toe areas. Do not wear high-heeled shoes, open-toed shoes, sandals, or shoes made of woven material. Tank tops, shorts or short skirts should not be worn in the laboratory. Disposable aprons are cheap and can effectively protect the body and clothing.
4. Confine long hair and loose clothing.
5. Always wash hands and arms with soap and water before, or immediately after, leaving the work area. This applies even if one has been wearing gloves.

6. Never perform any work when alone in the chemistry laboratory. At least two people must be present and one of those must be the instructor.
7. Never eat, drink, smoke or apply cosmetics in the chemistry laboratory. Fingernail biting or pencil chewing should be discouraged.
8. Never perform unauthorized work, preparations or experiments.
9. Never engage in horseplay, pranks, tricks, gags, or other acts of mischief in chemical work areas.
10. Never remove chemicals from the chemical work area without proper written authorization.

Rules appropriate to the local situation and the local laboratory should be added to this list whenever appropriate. This list is meant as a guideline and is not complete. Many laboratory books have suggested safety standards, more complete than this list, that you might wish to adopt.

General Guidelines and Hints

This section presents a collection of good ideas, expressions of common sense, guidelines or rules, and standards of action and behavior that are useful in maintaining a safe laboratory environment. These are not issued or dictated by any regulating agency but should be used to help establish the pleasant and safe working environment that the teacher will want in his or her area.

1. Visitors to the laboratory should be provided with eye protection (goggles), and, perhaps an apron or laboratory coat. It is just as important for visitors to use good safety practice as it is for students, teaching assistants or the instructor.
2. Every laboratory student should be visible from the teacher's normal vantage points. Blind spots are invitations for disaster.*(5)*
3. Students must be encouraged to report all accidents. In regrettable situations where classes have too many students, the teacher may not see an accident happen, and must rely on students for information. A student should never be embarrassed or penalized for identifying a safety problem.
4. Students, teachers and assistants must wear proper clothing and footwear in the laboratory. The feet are often as subject to injury as any other part of the body. Serious injury can result when reagents, glassware or apparatus fall on the floor and break or splash.
5. Never smell a compound directly. If required, one method by which a sample could be smelled indirectly is to cup the hand, pass the cupped palm over the container and toward the nose. Another method is to gently waft (fan) with your hand over the container in your direction.
6. It is important that spills and splashes of strong acidic and basic solutions be controlled and cleaned up immediately. When cleaning up acid or base spills, it is best to use absorbent materials. Safety supply companies offer "spill kits" for all sorts of liquid spills, including acids and bases. They are expensive. A suitable alternative is to have a convenient supply of salts distributed around the laboratory. These salts neutralize the spills and can be purchased in bulk at reasonable prices.

They are:

Sodium hydrogen carbonate (sodium bicarbonate), $NaHCO_3$, for **ACID** spills.

Sodium hydrogen sulfate (sodium bisulfate), $NaHSO_4$, for **BASE** spills.

7. The nature of concentrated sulfuric acid is a particular problem in the school laboratory. Concentrated sulfuric acid acts as a strong oxidizing agent and it reacts in an exothermic manner with many substances.

   Concentrated sulfuric acid also is a dehydrating agent by reaction with water. The heat of hydration is large enough to cause water to boil, leading to the adage: **"add acid to water, not water to acid**." It is this ability to dehydrate proteins and carbohydrates that results in their decomposition when in contact with sulfuric acid.

   Hot, concentrated sulfuric acid causes very serious thermal and chemical burns because it has a higher boiling point than water and a large heat capacity.

8. One should be aware that there is substantial difference in the possibility for injury between concentrated and dilute solutions of acids and bases. In virtually all cases, the concentrated form has far more potential risk than more dilute solutions. The concentrated solutions often have different chemical properties than more dilute forms. For instance, concentrated sulfuric acid, particularly when hot, is a strong oxidizing agent, while dilute sulfuric acid is not.

9. Students need to understand that "dilute" does not specify a relative degree of safety. Historically, dilute acids were acids of 6 M concentration which is, of course, much more concentrated than solutions of 1 M, 0.1 M, *etc.* Thus, a dilute solution of strong acid or base may have a potential for harm and must be treated with caution.

10. Special precautions must be taken when using strong bases. Eye protection is required as bases destroy eye tissues very rapidly. Strong bases also react with epidermal tissue (skin) and mucous tissue much more rapidly than acids. In addition, since pain is generally not immediately felt when most strong bases come in contact with the body, prolonged exposure often produces more severe burns than is the case for exposure to acids. With most strong acids, pain is felt nearly immediately.

11. Fire safety can be considerably enhanced if the teacher and school work toward avoiding the use of laboratory burners. If possible, replace laboratory burners with heating mantles, hot plates, sand baths or hot water baths.

12. Electrical outlets should have ground fault protectors and the instructor should be given ready access to the circuit breakers. These should be tested regularly.

13. The proper knowledge on the use and handling of fire extinguishers is important. The force of a blast from a fire extinguisher could be sufficient to knock glass containers off a table top, sometimes exacerbating a problem.

14. Two vital questions to ask before using any chemical are:

    How can this substance hurt someone?

    How can protection be offered against that possibility? *(5)*

15. Good ventilation in the laboratory, preparation room, and storage area is very important. Many chemicals are toxic but have no readily identifiable odor. In particular, since ingestion of aqueous solutions is unlikely, heavy metals are most dangerous when inhaled as dusts, aerosols or fumes.

16. Sunlight can have two possible negative consequences. Some chemicals decompose when exposed to light (for example silver nitrate is photosensitive); others may be sensitive to an increase in temperature produced by exposure to sunlight. Arrange the chemical storage area to prevent chemicals from being exposed to direct sunlight *(5)*.

17. Always put the purchase date on a container when you receive it, so that its age is known. Only fresh chemicals should be used, and those susceptible to decomposition should be properly disposed of when they have exceeded their useful lifetime. Some substances previously believed relatively safe may be susceptible to decomposition producing a potentially explosive mixture. (Some hydrocarbons, alcohols and ethers may produce explosive peroxides.) Less is better! Whenever possible, buy what you need in the amounts needed so that it is all used up. Do not store large amounts of any substance for long periods of time.

18. For safety reasons, the larger package of chemicals is generally not the best buy, although it may appear that way. The costs involved in storage, maintaining inventory, and waste disposal are significant.

19. If you transfer chemicals to a new container, be sure that the container is of a compatible material -- and be sure you transfer all the information on the label as well.

20. Waste chemicals should be stored in properly labeled containers identifying **all** the contents (especially in the case of heavy metal salts and chlorinated compounds). Waste disposal companies will not accept poorly labeled containers. Some of the wastes can be recycled (especially the heavy metals) if stored separately by type. It is not a good idea to have only one waste bottle for all the wastes generated. Reactions may occur creating a possible safety hazard. The practice of only one waste bottle also leads to expensive disposal costs.

21. Take just a few minutes each week to look around your classroom, laboratory, prep area and storeroom. Correct possible sources of accidents before they occur *(5)*.

22. Keep your storeroom locked. If a chemical is removed without your knowledge and an accident results, you could be held liable.

23. Keep a written record of all accidents, explaining the situation in detail. This record may help prevent accidents from reoccurring if it is reviewed periodically.

24. Some chemicals particularly damaging to the eyes are *(5)*:
    - alkalis and bases
    - dehydrating agents such as sulfuric acid and acetic anhydride
    - oxidizing agents such as potassium permanganate
    - any chemical which reacts with moisture
    - organic solvents

- phenols
- acids

25. Carbon disulfide has no place in a high school laboratory; it is a particularly hazardous chemical. It is poisonous and, due to its extremely low flash point, it can form an explosive mixture with oxygen over a very large composition range.

WHAT CAN YOU DO TO MINIMIZE RISK?

Risk is the possibility of suffering harm or loss. The degree of risk is a measure or estimate of the probability that the harm or loss will occur. Thus, the notions of risk are couched in the language of probability, statistics, projected analyses, estimates, reasonable guesses, and other such ill-defined terms. The process is essentially one of providing some quantitative estimate of potential harm.

Three truisms seem appropriate here:

1. No experiment is completely safe and without risk.
2. Any experiment worth doing is worth doing safely.
3. No concept is so important that it cannot be taught safely.

There are several fundamentally important steps that any instructor can take to reduce the degree of risk in any experiment or demonstration. Kaufman has developed and published a list of simple, low-cost actions that can make a significant impact on the degree of safety in the high school chemistry laboratory *(6)*. The listing for reference *(7)* includes information about publications available from the Laboratory Safety Workshop established by Kaufman.

1. Establish a school safety committee. The first function of that committee should be to do a needs assessment and safety audit, and then to itemize the various needs as a list of priorities. It is ultimately the responsibility of the committee to urge completion of those various tasks and to seek means by which those objectives can be accomplished.
2. Determine a list, and acquire books to make a departmental safety library that can be utilized when necessary. Then use the resources. Safety books are expensive and one must be judicious about the selections. There is a short bibliography in Appendix V that may help in this regard.
3. Maintain a list of material safety data sheets (MSDSs) in the science area. Manufacturers and suppliers are now required to provide an MSDS with each chemical sold, and it should be available to the end user: the teacher, instructor and laboratory worker. Each time a chemical is purchased by your school the supplier must send an MSDS. If you are not receiving MSDSs because they are being stored at some central store or administrative office, be sure to ask for copies so that you can have a set of MSDS documents in the science area.
4. Become familiar with the format of MSDS reports so that useful information can be obtained from them quickly when necessary.

5. Understand the safety requirements and attributes of every experiment you intend to use in the laboratory. Study the experiments to determine what special care should be taken and inform yourself and your students of those steps.

Common Sense

Any experiment can be either safe or dangerous, depending on how carefully the experimenter follows directions, the clarity of directions written for the procedure, how carefully an experiment is planned, the experience of the experimenter, how much attention is directed toward assuring safe conditions and other similar factors. Common sense is an important part of any safe experiment.

In the case of laboratory safety, at least, we are concerned with the following steps, all of which should be expected of a competent instructor:

1. Preparation and planning prior to the activity.
2. Presence in the laboratory during the activity.
3. Careful attention to student activity.
4. Thoughtful description of possible safety hazards.
5. Clear expression of means by which the activity may be safely completed to those performing the activity.

Teachers should always ask themselves:

- What can possibly go wrong?
- What can be done to prevent that event?
- What must I do if it does happen?

In the case of choosing an experiment that has some element of potential harm, it is useful to ask:

- Is this activity so educationally desirable that we can afford the risk?
- If there are hazardous chemicals required, can the same effect be demonstrated if less hazardous chemicals are substituted?
- Is there a better (safer) activity that can demonstrate the same principle, concept, technique, or method?

Since one cannot eliminate every activity that has an element of risk, *i.e.*, there are no activities that are *risk free*, the educator must make a conscious decision that the activity is sufficiently important to be included in the course syllabus.

Rules of Thumb

1. The instructor or laboratory assistant should pay attention to safe laboratory practice, and demonstrate such practice at every opportunity.
2. New techniques should be both described and demonstrated, and any safety precautions pointed out in the process.
3. The instructor must be a model of safe laboratory practice. It is not good enough to tell someone how to do it safely. This is particularly true with regard to protection such as eye goggles, gloves, laboratory apron or laboratory coat, *etc.*
4. The instructor must encourage students to be aware of correct and safe laboratory practice.
5. The instructor must choose activities whose educational value justifies the potential risk.
6. The instructor must prepare activities, experiments, written directions, and other materials on the assumption that aspects of safety should be taught and are legitimate areas for instruction.

The chemistry teacher must carefully evaluate the experiment or demonstration to be sure that:

1. *there is a reasonable justification for use of the chemicals,*
2. *the potential risks are understood,*
3. *less hazardous substitutes are not available,*
4. *the quantities to be used are as small as practical, and*
5. *the problems of waste disposal are within the capabilities of the teacher and the school* (2).

Classroom Management

Proper laboratory and stockroom management is important to minimize potential risks from chemicals and equipment. When chemicals are properly stored, correctly labeled, and presented in an orderly manner, the ability to find and use them effectively is greatly enhanced. This practice results in a savings of time and energy during the preparation phase of a laboratory exercise.

Classroom management, on the other hand has the added dimensions of finding a means by which one can effectively deal with the students in the class, and a means by which the subject matter can most easily be learned. These extra restrictions most certainly compound the problem of dealing with safety issues. Generally, good classroom management leads to a safer environment. It also saves time, effort, reduces tension, and other attributes that reduce our teaching effectiveness.

The key to good classroom management, and the development of a safe working environment, is the quality of planning that precedes the actual laboratory demonstration or experiment.

FACILITIES

General

There is a certain minimum of safety equipment that should be available in any chemical work area, including school laboratories. We recommend the following list of items be available: safety showers, eyewash fountains, (preferably with drains) fire extinguishers, adequate ventilation, fume hoods where required, wash sinks, and approved waste disposal facilities. Eyewash bottles are not suitable alternatives to eyewash fountains. They do not provide the flow necessary and can become contaminated.

All of these safety facilities should be conveniently located, present in sufficient quantity to provide useful service to the workers in the laboratory, be properly maintained, and tested on a regular basis.

The chemical work area should have appropriate facilities to allow quick response to potential emergencies. Such facilities, in addition to those described above, include the following:

1. A telephone with an outside line, strategically located and readily available, in working order. Emergency telephone numbers must be clearly posted nearby. A telephone line which goes through a school switchboard is often of no use after school hours. If an outside line requires internal access, be sure to provide all necessary access numbers. Suggested numbers include ambulance, poison control center, medical care, fire and police. In many areas these numbers may be accessed by dialing the 911 emergency response number.
2. Aisles and exits must be kept clear at all times.
3. There should be two or more well-marked and unobstructed exits for evacuating the laboratory.
4. A general alarm system, which alerts facility emergency and security services, should be provided. Automatic smoke and fire alarms are highly recommended and, in some cases, required by law, especially in the absence of automatic water sprinklers. These systems should be properly tested and maintained on a regular basis to ensure their proper operation and records should be kept.
5. Highly visible signs should be posted in areas where hazardous operations are being conducted or where toxic, reactive, or flammable chemicals are being used or stored.
6. Emergency drills are essential. In case of general evacuation, students, faculty and staff should be required to go to a designated area outside the building and remain there until accounted for.
7. Special consideration should be given to ensure ease of evacuation and accessibility to safety equipment by physically disadvantaged individuals.

The instructor should know, and have ready access to, the location of all safety shut-off devices, such as:

1. main natural gas valve for the entire laboratory area.
2. water shut off valves for the laboratory.

3. electrical circuit breakers.
4. switches or circuit breakers for hood fans.

Ventilation

High school laboratory experiments are generally conducted on laboratory benches and not in laboratory hoods. It is important to characterize the limits of available room ventilation in the laboratory so that the toxic materials, if any, in the atmosphere during an experiment can be assessed. If it is determined that OSHA toxic levels might be exceeded during a laboratory exercise, any of the following may be performed to reduce the potential hazard from the air:

1. Change the amounts of chemicals used in the experiments.
2. Reduce the number of students performing the experiment at any given time.
3. Substitute a less toxic chemical in the experiment.
4. Require certain experimental operations be done in the hood.

These requirements relate not only to the laboratory in which the demonstrations or experiments are carried out, but also to the preparation area and chemical storage areas, where the materials for the activity are prepared.

Periodic inspections of the ventilation system are required to assure good working conditions. The incoming supply of air must be adequate to meet the exhaust requirements, and there should be no circulation of exhaust air from hoods back into the laboratory or to any other part of the school facility.

Finally, the ventilation system of the chemical storage and preparation area, should be independent of any other air circulation system in the building.

Fume Hoods

One key to safe handling of chemicals is to use an enclosed, properly installed hood system. Saunders *(8)* addresses hood requirements in academic institutions:

> *Fume hoods must operate properly, and should be checked by the teacher on frequent, regular time intervals. Records must be kept of such testing.*
>
> *Experiments which produce flammable gas, toxic vapors, or noxious odors should be performed in hoods. It has been recommended that there should be a hood for each two workers, and each worker should have at least 2.5 linear feet of working space at the hood face.*
>
> *The exhaust rate is not a reliable single measure of hood performance, but air flow through the face of the hood should be somewhere in the range of 80-100 cubic feet per minute. Velometers should be used to survey hood characteristics to assure uniformity of air flow over the face of the hood.*

Fire Protection

Fire Extinguishers. Fire extinguishers must be posted both in the laboratory, and also in the preparation and storage areas. In most cases, carbon dioxide or halon fire extinguishers, suitable for Class A, B, and C fires should be used. They should be of sufficient capacity to provide a strong working condition over more than a minute total duration. Extinguishers should be placed in convenient locations, in unobstructed view, and near entrances to the laboratory or preparation and storage area. If the laboratory is very large, it may be desirable to place a second extinguisher in another area of the room.

The fire department or other contractor should check the condition of all fire extinguishers on a periodic basis. A suitable time interval is often taken to be 6 months to a year.

Every teacher should practice using a fire extinguisher so that it can be operated effectively in an emergency. When using a fire extinguisher, be sure that:

1. The force of the material emanating from the nozzle of the extinguisher may be sufficient to topple or push lightweight objects across or off a laboratory bench, possibly adding to the problem.
2. The flow from the extinguisher must be directed at the base of the flames in order to have the greatest effect. As hot air is produced in a fire its density is lowered and it rises, carrying the flames upward.
3. The sound of the extinguisher discharging may startle a first-time user. Exposure to the noise will prevent surprise at the time of a fire.
4. The duration of discharge of a typical fire extinguisher is quite short. It is important that the extinguisher resources are efficiently and effectively used and conserved.

Eyewash Fountains and Safety Showers

> *Eyewash fountains should provide a copious and gentle flow of temperate, aerated, potable water for a period of at least 15 minutes (15 minutes of cold water is intolerable). Plumbed installations are best, but the lack of a floor drain should not prohibit installation of the eyewash fountain. When possible, students should be encouraged to practice the procedure. Use of the hands should not be required to maintain the water flow.*
>
> *Eyewash fountains should be tested on a regular basis and a record kept of such tests. (The National Safety Council recommends a 3-minute flush of plumbed installations weekly to reduce the threat of eye infections.) ANSI Z358.1 requires that eyewash units should be located no more than 10 seconds in time nor more than 100 feet in distance from the area it serves. Their location should be clearly labeled (1).*

Showers. The location of safety showers must be clearly marked and readily accessible from all areas that they serve. ANSI Standard Z358.1 requires that showers be located no more than 10 seconds in time nor more than 100 feet in distance from the hazard or work area. Water flow must be sufficient to drench the subject rapidly and to accommodate more than one person. The valve should be of the pull on or pull off type. ANSI Z358.1 requires a minimum flow of 113.6 liters per minute (30 gallons per minute) of temperate, potable water. The shower should be tested on a regular basis. Records should be kept of testing.

Electrical Hazards

All electrical outlets must have a grounding connection (third prong), as required by OSHA rules. If equipment does not have such a cord and plug, it must be replaced. All electrical outlets should be protected by ground-fault interrupters, which do not work in two-wire installations.

Continuity of connections should be checked regularly by an authorized inspector and documented in a permanent record. The National Electrical Code should be followed in all installations. This includes proper grounding as well as proper equipment for the hazard area.

The instructor should check equipment to assure that frayed, worn, or stretched wiring is eliminated from the laboratory. Certainly, no wire should be stretched across the floor where someone could trip over it.

Definitions

Chemistry teachers are familiar with the classifications of chemicals that express the chemical reactions of those materials, but terms may be different when the safety or threat of potential hazard is incorporated into the definitions. Following are definitions of chemistry terms that include these safety considerations and the threat of potential hazard.

Acids and bases: Safety issues related to acids and bases are grounded on four concerns:

1. Many acids and bases produce chemical burns in contact with skin and other tissue. Thus, one must protect tissue from contact, and if contact is made, the skin should be washed with copious amounts of water. Strong bases often react faster and more severely with tissue than do acids, so their potential hazard is greater. Do not neutralize an acid or base spill on the body. A more severe burn may result.
2. Acids and bases react to produce large quantities of heat, particularly if they are strong acids and bases (defined next). That heat may cause thermal burns or other degradation.
3. The acid or base may react with active metals, producing hydrogen, which is potentially explosive.
4. Some acids are capable of acting as oxidizing agents, reacting with reductants to produce heat and, possibly, noxious gases or reactive products.

Strong acids or bases: A strong acid or base is one that is completely ionized in aqueous solutions. Strong acids and bases are likely to react faster and more vigorously than weak acids and bases and they may act as oxidizing agents. Thus, they are a greater threat than weaker acids and bases.

Oxidants (oxidizing agents, oxidizers): These are compounds that readily gain electrons, often contribute oxygen to the chemical reaction, and frequently react to produce heat, thereby increasing the rate of a reaction. Oxidants also initiate or promote a reaction with reductants which may lead to fires.

Reductants (reducing agents): Compounds that give up electrons. They react with oxidants and often serve as fuels.

Flash point: The minimum temperature at which a liquid gives off vapor in sufficient concentration to form an ignitable mixture with air.

Combustible liquid: A liquid having a flash point at or above 37.8°C (100°F).

Flammable liquid: A liquid having a flash point below 37.8°C (100°F).

Reactivity: A measure of how fast the substance will react, and the chemical nature of the reaction. Many safety applications limit reactivity to a consideration of whether it will react so quickly as to constitute an explosion. Other applications designate whether the substances are stable toward heat, pressure, or water. This term is widely misunderstood.

Explosion: A chemical reaction that is very rapid, produces heat and produces gases, which expand.

Carcinogen: A substance capable of causing cancer in humans.

A *Confirmed Human Carcinogen* is a substance that is recognized to have carcinogenic or cocarcinogenic potential in humans.

A *Suspect Human Carcinogen* is a substance suspected of inducing cancer based on human evidence or carcinogenesis in two or more animal species. Compounds having these properties, and those that are suspected of producing tumors in animals are often grouped together in warnings, so that some compounds may be designated in warnings as "carcinogenic" when, in fact, that property has not been shown for humans.

Mutagen: A substance capable of causing a change in the DNA in a human cell.

Teratogen: A substance capable of producing fetal malformations in humans.

Active Metal: Any metal that reacts with water, steam, or aqueous acid to produce hydrogen gas.

Corrosive: Causes chemical or thermal burns.

CHEMICALS

Incompatible Chemicals and Storage Considerations

Incompatible chemicals are those that will react together when accidentally mixed, producing large quantities of heat, explosive gases, fire, toxic or noxious gases. *It is because of reactivity among incompatible chemicals that schools are urged not to store chemicals on an alphabetical basis, but to arrange storage of chemicals according to reactivities.*

A brief summary of incompatible chemicals follows, along with an example of each. More elaborate lists of incompatible chemicals are available *(9)*.

1. **Acids and bases** should not be stored together; their reactions produce copious quantities of heat. In particular, sodium hydroxide should not be stored near sulfuric acid.
2. **Oxidants and reductants** should not be stored near each other. Mercury(II) nitrate (mercuric nitrate) and methyl alcohol (methanol, wood alcohol) react together to form mercury fulminate, an explosive hazard. Another example is the mixing of cleaning agents containing ammonia and a bleach containing hypochlorite ion (such as Purex®, Clorox® and other common bleaches), which may result in the production of chlorine, a poisonous gas or a chlorine oxide.
3. **Active metals and oxidizing agents** should not be stored together. An example is aluminum and bromine. Aluminum foil reacts with bromine at room temperature, producing sufficient energy to become incandescent.
4. **Volatile solvents and oxidizing agents** should not be stored near each other. These may spontaneously combust if a source of ignition is present, and some even burst into flames when a mechanical action, such as hammering, commences near it.
5. **Concentrated acetic acid** (glacial acetic acid) is of special concern. Not only is it an acid, but it is volatile, its odors are noxious, and may undergo other types of reactions. It is also a very flammable liquid.
6. **Concentrated nitric and concentrated perchloric acids** are other reagents of special concern, since they are strong acids, as well as strong oxidants, often producing toxic gases as products of their reactions. Perchloric acid is such a strong oxidant that we recommend it be removed from the school laboratory, if present.

Discussion of APPENDICES I and II

The High School Subcommittee of the ACS Joint Board/Council Committee on Chemical Safety is reluctant to act as censor of a list of chemicals which the high school instructor may wish to use in the laboratory. It is true, however, that there are some chemicals for which warnings are appropriate and for which comments may lead the instructor to choose to substitute some less hazardous substance, to discontinue use of this substance, or to store the chemical in a manner appropriate to its potential hazard.

The abbreviated list of chemicals in Appendix II is compiled from several sources, such as recommendations by state departments of education, publications from school districts with which we were familiar, and our own experience. They reflect both the suggestions of those sources, and modifications by the High School Subcommittee. One of those sources, for instance, suggests forbidding the use of all nitrate salts because nitrates are strong oxidants. In fact, aqueous solutions of nitrates, the form in which they are most often encountered in the laboratory, do not often act as oxidants.

Appendix I provides a list of chemicals that we suggest be removed from the chemical storeroom and not be used in high school chemistry and other science classes. We suggest that their use is not justified when the pedagogical value of their use is compared to the potential risk of their use. Appendix II offers **advice** as to the disposition of chemicals in the high school laboratory. These are not definitive lists, and do not represent anything more than an expression of our best experience with regard to their use in the high school laboratory. The teacher must be aware of their potential hazards, and try to minimize these whenever a chemical is chosen for use in the laboratory. While these lists will help in making that determination, the corresponding MSDSs should also be consulted.

Chemicals in Appendix II are identified by code letters as follows:

A. Those that are known carcinogens, mutagens, or teratogens. We generally do not recommend using these chemicals.

B. Those that are suspected carcinogens, mutagens, or teratogens. We recommend use of these chemicals be greatly restricted.

C. Those found on Flinn's list of "40 Devils" *(5)*, which represent routine sources of trouble in the high school laboratory. Use of these chemicals should also be avoided.

D. Those that pose sufficient risk, other than those listed above, that they probably should not be in the high school laboratory. We generally do not recommend using these chemicals.

E. Those that may be used by a qualified chemistry teacher, but probably should not be used by students.

F. Those that have potential risk associated with them, but whose use by students may be justified, providing the smallest amounts possible are used and the students have been informed as to the correct means of handling these chemicals.

Thus, prudent use of most of the chemicals on this list has been recognized.

The Handling and Disposal of Chemicals

Once the teacher has chosen to use a particular chemical during an experiment, then two issues remain relative to "handling." First, "what can I do to reduce potential hazards?" And second, "what is the best means by which the products of the reaction or the chemicals used during the experiment may be disposed?"

There is not sufficient space here to offer a procedure for each chemical that one may choose to use, but some general guidelines can be provided.

1. Suppliers of chemicals will indicate on the MSDS whether the chemical is stable over time, and what the anticipated shelf life is. This is the best source of information concerning the storage of chemicals. Such information is also available in various sources, such as Flinn's catalog.
2. Aqueous solutions of most salts are relatively stable and, if properly labeled, may be stored indefinitely.

3. Acids and bases may be discarded into the sink after they have been neutralized and they should be washed down the drain with large amounts of water. Although this procedure is discouraged, the direct disposal of waste acids or bases to the sink should be done slowly, with mixing, and with flowing water. It is always easier to dispose of more dilute solutions and best to dispose of properly neutralized solutions. Alternatively, more concentrated acid or base solutions could be properly labeled and stored, to later serve as a source of that acid or base in a more dilute form.

4. Strong oxidants can be effectively reduced by using sodium hydrogen sulfite, $NaHSO_3$, in aqueous solution. Treatment must be done in a hood or well ventilated area since small amounts of sulfur dioxide may be produced if the solution is acidic.

5. Most secondary schools require such small amounts of organic solvents for a given experiment that the waste solvents may be disposed of by placing an open container in a hood and allowing the solvents to evaporate into the air. The door of the hood should be closed so that vapors do not escape into the room during this process.

6. Procedures for dealing with the disposal of salts of heavy metals in aqueous solution including (but not limited to) antimony(II), barium(II), bismuth(III), cadmium(II), chromium(III), cobalt(II), copper(II), lead(II), manganese(II), nickel(II), silver(I) and tin(II) are suggested below. All of the metals commonly found in high school laboratories precipitate as either metal oxides or hydroxides. Most can be precipitated with 1 M sodium hydroxide (NaOH, pH 14). Aluminum, lead(II), tin(II), tin(IV), and zinc(II) are the few metal hydroxides or oxides common in high schools that redissolve at high pH's. To precipitate these ions, the pH must be maintained between 7 and 8. A more complete discussion of precipitation of metals with sodium hydroxide, or those that must be precipitated by sulfides, can be found in *Prudent Practices for Disposal of Chemicals from Laboratories (10)*.

   a) Most metals, and especially the valuable ones such as silver and mercury (from broken thermometers), should be collected separately, stored and recycled (the new three r's: recover, recycle, reuse). There are many companies that specialize in reclaiming metals and will take the waste metals. The original supplier should be contacted for information on names and locations of reprocessors *(10)*. Reclamation procedures for a number of metals have appeared in articles in the *Journal of Chemical Education* (*Appendix* VI).

   b) Although we urge that most metals be recovered, recycled and reused, metal hydroxides that will be discarded can be filtered, dried, collected, labeled, and transported to an approved dump site. The dried salts can be stored together. A mixture of salts is not inappropriate so long as each of the metals in the mixture is properly identified on the label of the container. Since the solids require much less space than the dilute aqueous solutions most often encountered in the laboratory, residue from several years' experiments could be collected until the volume is suitable for disposal. The disposal of chemicals is increasingly more difficult. Regulations concerning disposal of chemical wastes vary from state to state, however it is generally true that disposal companies will not accept unlabeled, or improperly labeled, wastes.

CHEMICAL STORAGE AND LABELING

Inventory Management

Wahl *(2)* introduces the characteristics of modern laboratory management with the following statement:

> *A comprehensive chemical management system is an essential element of any science program, and particularly of a high school chemistry program. With increased federal, state and local regulations and ever more frequent litigation, the need for such management systems becomes even more obvious.*

A good chemical management system is useful from several points of view. It fosters useful economic practices in the laboratory, it saves precious time, it fosters good storage habits and reduces potential storage problems of old chemicals, it reduces potential legal challenges, and it fosters good educational practices. A good inventory system will have the following effects:

1. Reduction in costs.
2. Timely detection of missing or stolen chemicals.
3. Control of useful "lifetimes" of chemicals can be monitored. In this manner, chemicals can be used before their "expiration date," and those that exceed their shelf life will be discarded since they will be clearly identified.
4. Overstocking of chemicals is avoided.
5. The potential for chemical accidents is reduced since proper storage habits are constantly being monitored.
6. Provides written documentation of good and prudent practice in case of litigation.
7. It helps develop the educational practice of making thoughtful and conscious decisions as to what work will be supported in the laboratory.
8. It will set a model of good practices for students and other school personnel.
9. It reduces the problem of waste disposal, since unwanted chemicals are not present.

> *It is important for the teacher that an inventory of chemicals, both reagents and laboratory solutions, be done, maintained, and kept up-to-date. Such an inventory should meet the requirements of local safety and health codes, fire department regulations,* etc.
>
> *Such an inventory will provide the benefits shown above, and will apprise the instructor of any hazardous chemicals that should be disposed of or which require special permits.*
>
> *The inventory will allow one to assess which chemicals are not or have not been used, will provide some information relative to deterioration, and will provide an opportunity to relabel those bottles for which the label is not suitable (6a).*

Characteristics of a Good Inventory System

The previous section justified the establishment of a good laboratory management and chemical inventory system. What kinds of information and practices should be a part of such a system?

1. General information. The inventory list should clearly identify the school, the room being inventoried, the address and telephone number, and the name of the person in charge. It should have the date of the inventory (and a new date each time a new revision is produced).
2. It should have a listing of each chemical, including the following information:
   a) Location of the chemical. Storage will be discussed in a later section.
   b) Cross-listing of trade or common names.
   c) Chemical formula or description of form (crystal, powder, bar, *etc.*).
   d) Quantity.
   e) Receipt or purchase date. This date should also be written on the container the day it is received.
   f) Expiration Date. This date should also be listed on the container.

Shelf Life

Many chemicals deteriorate with storage time. Manufacturers are encouraged to provide the date of packaging and analysis of each chemical [that is] distributed. Materials used to package chemicals may also deteriorate and allow the contents to react with the atmosphere. As the quality of the chemical changes with age, there will be a tendency not to use old chemicals in experimental work.

Chemical manufacturers occasionally list an 'expiration date' for a given chemical, such as isopropyl ether, and indicate specific storage precautions. A definite date is stamped on the bottle or can and precautions and disposal instructions are provided. There is really no help, however, provided to establish the 'shelf life.'

[Teachers] must recognize that many chemicals cannot be stored indefinitely and that storage guidelines must be established. For disposal purposes, 'decision dates' are more useful than 'expiration' dates. When stored chemicals have reached the end of the allocated time span, [or can be seen to have deteriorated], they should be removed and disposed by an appropriate means.. Visual inspection is important in the disposal decision. Indications for disposal of chemicals include the following:

- slightly cloudy liquids
- darkening or color changes
- spotting on solids
- caking of anhydrous materials
- existence of solids in liquids (or liquids in solids)
- pressure build-up in bottles
- evidences of reaction with water
- strange or unexpected odors
- damaged container

Storage

Flinn *(5)* has suggested, and others have adopted, an organization of chemicals useful in the high school storeroom and laboratory. The system is designed to isolate incompatible chemicals from each other, and reduce the number of possible chemical interactions to as few as possible. When one looks at inorganic chemicals, in particular, one finds that the potential hazards are often due to the anion part of the salt, and it is on the basis of anion reactivities that the suggested system has been developed.

Flinn's suggested groups are as follows:

1.	Inorganic (I1)	Metals, metal hydrides [must be stored away from water]
2.	Inorganic (I2)	Acetates, halides, sulfates, sulfites, thiosulfates, phosphates, halogens, *etc.*
3.	Inorganic (I3)	Amides, nitrates, nitrites, azides [but not ammonium nitrate]
4.	Inorganic (I4)	Hydroxides, oxides, silicates, carbonates, carbon
5.	Inorganic (I5)	Sulfides, selenides, phosphides, carbides, nitrides
6.	Inorganic (I6)	Chlorates, perchlorates, perchloric acid, chlorites, hypochlorites, peroxides, hydrogen peroxide
7.	Inorganic (I7)	Arsenates, cyanides, cyanates
8.	Inorganic (I8)	Borates, chromates, manganates, permanganates
9.	Inorganic (I9)	Acids [except nitric acid]
10.	Inorganic (I10)	Sulfur, phosphorus, arsenic, phosphorus oxides
11.	Organic (O1)	Acids, anhydrides, peracids
12.	Organic (O2)	Alcohols, glycols, amines, amides, imines
13.	Organic (O3)	Hydrocarbons, esters, aldehydes
14.	Organic (O4)	Ethers, ketones, ketenes, halogenated hydrocarbons, ethylene oxide
15.	Organic (O5)	Epoxy compounds, isocyanates
16.	Organic (O6)	Peroxides, hydroperoxides, azides
17.	Organic (O7)	Sulfides, polysulfides, sulfoxides, nitriles
18.	Organic (O8)	Phenols, cresols
19.	Misc. (M1)	Nitric acid

[**Note**: These are Flinn's suggestions. They are not complete, and Flinn Scientific Co. should not be expected to carry any responsibility. The instructor and school system must decide how to organize their inventory. It is true, however, that this system makes more sense than simple alphabetical ordering of chemicals.]

Overall storage management involves limiting the possibilities for chemical interactions under any circumstances. Once the inventory has been completed and the chemicals have been arranged according to the predetermined shelf storage plan, the storage area must be secured.

Inventory lists should be available in the storage area, the teacher's files, and with the school administration. Recent studies in North Carolina *(2)* suggest that a typical school may possess about 500 different chemicals.

Labeling

As a chemical is received into the school laboratory it should have a label which identifies its name, provides certain other information, but the different manufacturers have such different labels that one cannot make generalizations beyond that. Some chemical suppliers are making an effort to provide safety data both on their labels (Flinn, Fisher, J.T. Baker) and in their catalogs (Flinn, Fisher, Baker, Aldrich). Other sources of information about hazardous materials include the publications listed in the bibliographic section, Laboratory Safety Supply Co. (whose address is provided in a later section), the National Fire Protection Association, and others.

The label serves two purposes relative to safety. It should provide information to the person responsible for storage, safety, and maintenance of the chemical inventory, and it should provide information about the chemical properties of the substance(s) in the bottle. The appropriate standard for labels is given by the ANSI Z1290.1, *Standard for Hazardous Industrial Chemicals–Precautionary Labeling (4).* Requirements established by your state should also be consulted.

The **minimum** information that should be shown on each label includes the following:

1. Name (IUPAC Name)
2. Cross-referenced name(s)
3. Formula
4. Date purchased
5. Decision date (for disposal)
6. Degree of hazard (extreme, danger, warning, caution)
7. Type of hazard(s) (carcinogen, teratogen, flammable, poison, corrosive, toxic, explosive, irritant) and part(s) affected (skin, eyes, respiratory tract, *etc.*).
8. Lachrymator or harmful vapor
9. Precautionary measures (keep away from heat, *etc.*)

The NFPA (National Fire Protection Association) has developed a labeling system for chemicals that specifies hazards on the basis of:

health (*blue*, ranging from **deadly** [4] to **normal material** [0]),

fire hazard (*red*, ranging from **flash point below 37.8°C** [4] to **will not burn** [0]),

reactivity (*yellow*, ranging from **may detonate** [4] to **stable** [0]), and

specific hazard (*white*, including **acid**, **alkali**, **corrosive**, **oxidizer**, **polymerizes**, **radioactive**, and **use no water**).

Other companies, such as Flinn *(5)* and J.T. Baker *(11)* have developed similar systems, which may be of more use in the school laboratory since the NFPA list is somewhat limited and is oriented to those chemicals that are likely to burn, cause fires, or explode. One problem that is yet to be resolved is that there is no standard for color or symbol codes; those used by NFPA may not be the same as those developed by other companies or agencies. The school should decide which system(s) to use and how they will be used, and make that information known to all interested persons. Safety systems, symbols and pictograms are discussed in more detail in Appendix III (pp. 57-60).

SAFETY EQUIPMENT

Personal Protection

No one (teacher, visitor, student, helper, *etc.*) should work in the chemistry laboratory without certain items of personal protection. Each person in the laboratory, preparation, storage, or disposal area should have a pair of shoes with closed toes, goggles (see next section), and clothing to cover legs and arms.

In addition, other items may be used as required by the particular operation, and should be available. These include such items as full face shields, gloves (the kind depending on the use), laboratory coat, apron, respirator, *etc.*

Eye Protection

There are many types of eye protection useful to humans, but the number is limited for use in the chemistry laboratory. General recommendations for eye and face protection are provided in ANSI Z87.1.

The most common form of eye protection is a pair of goggles, whose design should meet the standards specified above. No one should work in the chemistry laboratory without appropriate goggles. The goggles should protect the eyes from irritating mists, chemical splashes and flying debris or projectiles (metal or glass). The goggles should protect a person from the front, top, and sides. Thus, the goggles should fit comfortably against the face, either be flexible or have a cushioned fitting to the face, have a window of unbreakable plastic, and have appropriate ventilation which inhibits liquids leaking through the ventilation openings. The goggles should be resistant to acids, bases, and organic solvents.

Full face shields should be used when there is a possibility of high temperature exposure, splashes from molten metals or salts, or the presence of hot sparks.

> *In the chemical laboratory, face shields alone are not considered adequate eye protection and must be used in conjunction with other eye protection. Goggles are recommended.*
>
> *Normally, contact lenses should not be worn in the chemical laboratory or workplace and they are never a substitute for eye protection. If contact lenses are permitted, goggles must also be worn at all times. Gases and vapors can concentrate under the lenses and cause permanent eye damage. Soft contact lenses may pose even greater hazard than hard contact lenses in this regard.*

> *Contact lenses can also trap particulate foreign matter in the eye and thereby produce abrasion of the cornea.*
>
> *In the event of a chemical splash into an eye, capillary action tends to hold the offending substance under the contact lens and against the surface of the cornea. At the same time, removal of the contact lens to achieve immediate irrigation is made nearly impossible by involuntary spasm of the eyelid. Persons attempting to irrigate the eyes of an unconscious victim may be unaware that contact lenses are in place, thereby reducing the effectiveness of the wash (1).*

We strongly urge students, teachers, instructor and assistants not to wear contact lenses in the chemical laboratory, and also urge that the teacher, instructor and any helper be aware of any student who persists in wearing contact lenses. A label should be placed on the goggles of those persons who also wear contact lenses in case of an emergency.

Clothing

> *Aprons, laboratory coats, gloves and other protective clothing, preferably made of chemically inert material, should be readily available and used.Gloves must provide sufficient arm protection to minimize the chance of spilled chemicals making contact with the skin. The selection of proper glove material is important (1).*

Gloves

Gloves common in the school laboratory are generally made of the following materials: polyethylene (clear, light-weight, flexible plastic); neoprene (black rubber); nitrile (flexible rubber, household rubber gloves); and latex (thin rubber, flexible, stretchy, surgeon's gloves).

Gloves of each of those four major types should be available for the instructor and workers in the storage, preparation and disposal area. Polyethylene gloves should be available for general student use. These gloves are inexpensive, disposable, and each glove can be used on either hand. In most cases for students, a wrist-length glove is suitable.

Housekeeping

Laboratory, storage and preparation areas are subject to the "Prudent Person Rule," which suggests that if a prudent person can reasonably have predicted the response at an accident site, then that person must assume the consequences of allowing the situation. Thus, when an instructor or laboratory helper leaves a cluttered desk or "mess" on the table, and the prudent person could have predicted such a "mess" would lead to an accident, that instructor or laboratory helper must assume responsibility for the results. Experience convinces the prudent person that the chance of an accident increases proportionately to the degree of disorder in the work area.

One of the easiest procedures that can be done to reduce potential risk and the chance of an accident is to keep the work area clean. Several additional steps will help in this regard.

1. Shelves on which chemicals are stored should have a 2-inch lip, so that bottles will not slide or roll off the shelf onto the floor or chemicals stored below.
2. No chemicals should be stored on the floor, on top or above the laboratory bench.
3. Chemicals should immediately be placed in appropriate containers: storage (if they are to be kept) or waste (if they are headed toward disposal).
4. Chemical spills should be **immediately** cleaned up.
5. Personal items (purses, clothing, books, papers, back packs, *etc.*) should not be placed in the vicinity of the chemical work area.

Transporting Chemicals

Once a chemical has been inventoried, cataloged, and stored, transportation or movement of the chemical from one place to another provides an area of concern. The degree of concern can be significantly reduced by following a few simple precautions.

1. Purchase chemicals in small containers. Even though the cost per unit may be slightly higher, the reduced risk of an accident provides ample justification for using several small containers rather than a large one.
2. Carry or move as small an amount as possible.
3. When transporting compressed gases in cylinders, use an acceptable gas cylinder cart on which the tank can be chained into position and firmly attached to the cart. The valve cover must be in place.
4. When carrying pint-sized or larger bottles of liquid reagent, do so in a flexible rubber or polypropylene bucket. The latter are even available with lids. Should the bucket be dropped or the bottle otherwise broken, the contents will be contained in the bucket, thereby limiting a dangerous spill. When possible, purchase chemicals in plastic coated bottles.
5. Transport bottles of chemicals in trays with lips, so that spills are confined to a small area. This is a reasonable rule for either solids or liquids, pure or mixed.

Unattended Experiments

No experiment, whether performed by a student or teacher, whether as an experiment or demonstration, should be left unattended. No student(s) should be allowed to work in the laboratory without the teacher or other qualified supervisor present.

SAFETY TRAINING

> *Training with regard to safe practices in the laboratory is the direct responsibility of the academic institution or school district, and is required by law. The agency (school district, school, institution, etc.) is responsible for training all employees. In the chemical laboratory, the institution is responsible for training not only the students who are taking the various laboratory courses, but also the assistants provided by the department. Students are expected to adhere to all safety rules and to participate conscientiously in any training exercises. Furthermore, students have the responsibility to seek advice and guidance whenever they are in doubt about safety procedures or potential hazards in their laboratory work (1).*

Communicating Hazards

> *Instruction should be given to students and staff regarding hazards of the chemicals being used in a particular course, and the manner in which these chemicals are to be handled and disposed of safely. The information provided by the various references, on the labels, and on material safety data sheets (which are now legally required) should be [clearly] understood [before usage] (1).*

Safety training is an ongoing process. Although the instructor should spend time early in the course to advise students as to the policies of the laboratory and school, the instructor should make an effort to cite safety precautions related to each laboratory experiment or activity at the time they are to be performed. Instructions and warnings could be written or oral, or both, depending on the situation.

Responsibility

The purpose for safety activities and documents such as this one is to foster safe activities that are educationally sound and which communicate the nature of science to students. The objective of safety rules and guidelines is to promote a safe environment in which that learning may take place. Laboratory is important to the communication of science concepts and understanding.

Safety guidelines are not promulgated solely in order to reduce the risk of some potential legal action, although that may be the natural outcome of maintaining a safe environment in which effective instruction can take place.

The school district has a responsibility to its employees (including teachers and laboratory helpers) to assure that they are aware of the appropriate safety practice and policy. The institution must post its safety policy in the laboratory area for the information of any person who might be in that area.

Instructors, teachers, and others working with students in the laboratory have the responsibility of informing students of potential sources of risk in the particular activity at hand.

Evacuation and Fire Drills

> *Each student should be aware of the location of the fire exits and routes to be used during an emergency, alarms (and how to use them), and telephones available during regular school hours, as well as after hours. Instruction in fire drills should be scheduled on a regular basis (1).*

Each exit from the laboratory should have a map posted giving directions for egress from the room in the case of emergency.

Fire Extinguishers

> *Each student and staff member associated with a particular laboratory area should be aware of the location of fire extinguishers and should be trained annually in the proper method of operating them. No one without the necessary "hands-on" experience of training should operate a fire extinguisher (1).*

Fires and combustible materials are classified into one of four groups, identified by a letter, in order to specify the type of fire extinguishers suitable for a particular material. Carbon dioxide, for instance, is a suitable extinguisher material for burning paper, but not for burning magnesium metal, since magnesium burns in a carbon dioxide atmosphere.

- Class A fires are fueled by paper and combustible solids.
- Class B fires are fueled by flammable or combustible liquids.
- Class C fires are electrical fires.
- Class D fires are fires resulting from burning metals.

Class D - Extinguishers suitable for this class of fire are not suitable for any other class; they are very specialized because of the very high temperatures produced by burning metals, such as magnesium, sodium, aluminum, *etc.* Class D fires will not be extinguished by water, foam, carbon dioxide, or halon.

Class C - Water must not be used on Class B, C, or D fires, and therefore isn't of much use in the chemical laboratory. Carbon dioxide, halon, or dry chemical extinguishers (either the sodium carbonate or monoammonium phosphate extinguishers) are suitable for Class C electrical fires. Dry chemical extinguishers have the important disadvantage that they present severe cleanup problems, which are especially troublesome around electronic equipment.

Class B - Probably the most important potential source of fire in the chemistry laboratory are flammable solvents. Halon, carbon dioxide and dry chemical extinguishers are suitable for this class of fire. Again, note the strong disadvantage of the dry chemical fire extinguishers. Halon extinguishers are probably the best overall choice, but they are more expensive than carbon dioxide.

Class A - Water may be used for these fires, carbon dioxide is less useful, and halon is probably the extinguisher of choice.

A review of the data above suggests that if you can have only one fire extinguisher, it ought to be a high capacity halon extinguisher. If you can have two or more, a carbon dioxide extinguisher would probably be the second choice, since it is a good general purpose extinguisher that is comparatively inexpensive.

Emergency Response

Listed here are minimum suggestions, which may vary according to local conditions and laws.

1. The chemistry laboratory must have policies and routes of egress clearly posted.
2. The laboratory should be prepared for emergencies by having some rapid means of contacting the proper authorities. It is suggested that the chemistry laboratory should have a telephone with an outside line, strategically located and readily available, in working order. Emergency telephone numbers must be clearly posted nearby. A telephone line which goes through a school switchboard is often of no use after school hours. If an outside line requires internal access, be sure to provide all necessary access numbers. Suggested numbers include ambulance, poison control center, medical care, fire and police. In many areas these numbers may be accessed by dialing the 911 emergency response number. Telephone numbers to the school administrative office and school nurse should also be posted.
3. Some schools provide other forms of intercom or communications networking, and the instructor and others responsible in the chemistry laboratory, preparation, storage and disposal areas must be familiar with the mechanisms by which emergency responses are to be handled.

It is the institution's or district's responsibility to see that all laboratory personnel and employees that may be in the laboratory are trained to follow the school's guidelines for emergency response.

Accident Reporting

> *All accidents should be reported. Accidents resulting in even minor medical treatment or observation must be recorded. A formal, written report stating the causes and consequences of each accident should be made to the designated authorities, including the insurance carrier. Recommendations for the prevention of recurrences should be discussed with the safety or risk officer, but also by the school's safety committee.*
>
> *A written report of each accident should be submitted to the administration and to the school's safety committee. They should take appropriate action to prevent recurrence. A periodic review of accident reports will often reveal problem areas that need special attention. There are certain OSHA reporting requirements which may apply depending on the institution and the nature of the accident. Unusual or unexplainable chemical accidents should be reported (1).*

Material Data Safety Sheets

Material safety data sheets (MSDSs) are safety reports that are supplied with each chemical purchased. The supplier or manufacturer is required to provide an MSDS with each chemical. School districts that have a central receiving area may keep MSDSs in a central location, but they must be available to any employee that uses those chemicals at any time. Thus, the chemistry instructor should maintain a MSDS for each chemical in the inventory on file. **Note:** currently, there is no standardized form of MSDSs, and the information in some cases has not been verified for accuracy, therefore at present we recommend using established reference texts such as *The Merck Index.*

The MSDSs cite threshold limit values (TLV) and permissible exposure limits (PEL) data for chemicals, they offer suggestions for warnings or precautions, such as whether the chemical can be absorbed through intact skin, and they suggest appropriate protective wear. Threshold limit values refer to "the average concentration of a chemical that is thought most people can be exposed to for a working lifetime with no ill effects." *(1)* The PEL is a legal standard issued by OSHA and is often identical to the TLV for a given chemical *(1).*

The employer must provide training so that employees know how to read and understand the MSDS. If students are employed in the science laboratories, they are also subject to an appropriate training program. The following information, taken from the booklet *Safety in the Academic Chemical Laboratory,* although voluminous, is pertinent to the use and understanding of MSDS information.

> *Material safety data sheets for hazardous chemicals and mixtures of hazardous chemicals cite the TLV's or PEL's for the subject chemicals or mixtures. The safety officer should ensure by appropriate measurements of the breathing air that all persons who work with hazardous chemicals are exposed at concentrations less than the TLV or PEL. Several methods are available to make the appropriate measurements and to suggest ways to improve the ventilation. Of course, a MSDS cannot cite a TLV or PEL if the limit has not been established. If available, the results of animal testing may provide some guidance. In rare cases, the chemical is not harmful to inhale and no limit is necessary. More often, the chemical will be harmful when inhaled and laboratory personnel must be protected.*
>
> *A MSDS also describes other hazardous characteristics of a chemical or mixture. Thus, if a chemical can be absorbed through the intact skin, the MSDS will so state and will also prescribe the use of protective equipment, such as gloves or full protective clothing, as appropriate. If the MSDS does not describe the material (e.g., rubber, neoprene, or polyethylene, for gloves) the supplier may be able to suggest a suitable material.*
>
> *. . . known and anticipated carcinogens recognized as such by competent authorities must be identified in the MSDS. But, there is as yet no corresponding recognized basis of authority for the identification of mutagens (substances capable of causing permanent, transmissible alterations in the genetic coding) or teratogens (agents interfering with normal prenatal development causing abnormalities in the fetus). Hence, a MSDS may be silent concerning mutagenic or teratogenic hazard due to lack of information.*

A variety of allergens (agents that can produce an immunologic reaction) may be encountered in the laboratory when suitable precautions are not enforced. Further, some chemicals evoke nonimmunologic reactions in individuals who are sensitive to them. Typical responses include dermatitis or asthma. The special problem is one of sensitization. Difficulties can also arise because the allergic response may not be readily identifiable. There is usually no physical reaction at the time of initial exposure although this is the point at which the sensitization may have occurred. The physical reaction then takes place after the next, or some subsequent exposure. Unfortunately, no common listing is available for compounds observed to act as allergens. Refer to NIOSH allergenic lists and the appropriate MSDS.

There are many factors influencing the toxicity of chemicals, some known, some not yet fully understood. Hence all chemicals should be handled with respect for their known or potential hazards. Skin, eyes and respiratory tract should always be protected from possible exposure by use of engineering controls, appropriate protective equipment, ventilation controls, and safe work practices, including personal hygiene practices. Eating, drinking, smoking and the application of cosmetics should be strictly prohibited in all areas where chemicals are handled. Thorough washing with soap and water should be mandatory at the end of a laboratory period and before leaving the work area during a break. These precautions can prevent many undesirable toxic consequences.

As required by the Hazard Communication Standard, an OSHA regulation found in 29 CFR 1910.1200, (10) and as required by similar regulations in most states, MSDSs are references to be used principally for the training of workers concerning the hazards and precautionary measures applicable to those particular chemicals that workers will handle in the workplace.

OSHA also requires that, for every hazardous chemical on the premises, a MSDS be readily available for employees' perusal. Because there is some level of hazard associated with every chemical, OSHA generally interprets all chemicals as hazardous. OSHA imposes no restriction on the sources of the MSDSs that are made available to the workers, but it does require all suppliers of hazardous chemicals to furnish copies of MSDSs to their customers. Therefore, this is the usual source of MSDSs for the employers' workers. In some states, the MSDSs must be from the supplier to fulfill the state legal requirements.

OSHA has prepared a 'suggested' format for a MSDS, but any format that supplies the OSHA required information is acceptable. This includes information stored in a computer database provided that the information can be accessed readily. The principal parts required in a MSDS include physical data, fire and explosion hazards, toxicity hazards, other health hazards, propensity to react vigorously (often called an 'incompatible chemicals' or 'reactivity' list), spill and leak clean up procedures, and above all, precautionary measures which, if taken, will materially reduce the probability of harm when the chemical is used. Therefore, when properly prepared, a MSDS is a useful tool in the training of users, including students who will use that chemical in their laboratory work. Note that although OSHA has no authority to require the training of students

> *concerning hazardous chemicals, faculty can take advantage of the availability of the MSDSs that are furnished by chemical suppliers and use these in the training of students concerning chemical safety and health hazards and precautions.*
>
> *A MSDS is always a tertiary source of information and not all MSDSs are prepared from reliable sources. Therefore, some MSDSs are less accurately informative than others. Not all fully describe the known hazardous properties of the substance and some describe hazards that are not attributable to that chemical. It is often useful to compare MSDS data from different suppliers for the same chemical. Better yet, refer to the primary literature; or, because this is often difficult to obtain, refer to reliable review sources, such as* **Handling Chemicals Safely** *(12), to the NIOSH / OSHA Occupational Health Guidelines (13), or to the books by Patty (14), Gleason (15), and Bretherick (9, 16), or the NIOSH Criteria Documents.*

[**Note**: We recommend you look first to those references found in *Appendix IV* of this document.]

> *MSDSs can provide unique opportunities for students to gain insight into how practical decisions are made when the use of a chemical is being considered. MSDS sections on toxicity and precautions guide faculty in determining the selection of one chemical for student use in place of another that, chemically, would 'work' well in a particular experiment. For example, the substitution of slightly toxic strontium chloride for the very toxic barium chloride in a laboratory study of insoluble sulfates. Similarly, a MSDS section on the reactivity of acetic acid with potassium permanganate or any other oxidizer reminds faculty to limit the amounts to be used by students in an experimental investigation.*
>
> *MSDSs are as reliable as it is possible to obtain, or produce locally, should be available to students, teachers and staff. Hazard and precautionary information from MSDSs should be incorporated as part of all laboratory procedures. Student laboratory reports should include safety and health concerns, the management of hazardous situations, and where appropriate, the proper disposal of both products and left-over reagents. Students should be encouraged to understand that the practical information found in the reactivity section of a MSDS is very important in designing safe experimental procedures. It is important for academic institutions [including high schools] to involve MSDSs in all aspects of laboratory work. It is equally important to instruct students so that later, after graduation, they and those whom they supervise will handle chemicals safely (1).*

FIRST AID

We urge all teachers to prepare themselves by attending a first aid training course, such as those provided by the American Red Cross, the local fire department, or other similar agency. We believe appropriate training in CPR is particularly useful for the teaching of chemistry in a safe manner, and teachers should make an effort, supported by the school district or supervisor, to acquire such training and to update it at appropriate intervals. Pamphlets regarding first aid can be obtained from the American Red Cross or the local fire department.

STUDENTS WITH PHYSICAL DISABILITIES

There are both state and federal regulations governing the provision for handling persons with physical disabilities in schools and other institutions. Because of its experimental nature, chemistry instruction offers special challenges when the class contains these persons. Yet experience has demonstrated, and continues to do so, that disabled persons can be taught chemistry, and they can perform most of the same experiments in the laboratory that are required of other students. Most experiments can be used as is, or suitably adapted for the physically disabled (See *Chemsource* module on *Teaching Chemistry to Persons with Disabilities*).

Instructors who face this problem must be given appropriate assistance from the school, principal, supervisor, or other persons of authority. For instance, the preparation of experiments for the handicapped person may require substantial expenditures of time over and above that required for the more normal experiment; thus, the school administrator must make provision to provide time compensation for the instructor.

In some cases a special area must be provided, but modern techniques, such as small-scale laboratories, and the possible availability of another student to work with the disabled individual may produce a very worthwhile instructional experience.

Other requirements and exemptions must be made in accordance with the local regulations. There is material available from the ACS Committee on the Chemists with Disabilities.

WASTE DISPOSAL

General Comments

Because of the interface between regulations formulated by state agencies, the EPA, the Department of Transportation, and other agencies, waste disposal has become the principal concern for typical high school teachers. It is not possible to anticipate every situation in which instructors may find themselves with regard to disposal of unwanted or hazardous chemicals, but certain general guidelines are appropriate.

To our knowledge, the best single, convenient guideline for disposal of common chemicals from the school laboratory is that provided by the Flinn Scientific Co. *(5)*.

1. Acids should be neutralized with sodium hydrogen carbonate ($NaHCO_3$, also known as sodium bicarbonate) before being washed down the sink. In particular, hydrofluoric (HF) and perchloric acid ($HClO_4$) pose special disposal problems.
2. Hydrofluoric acid (HF) *must* be neutralized before being poured down the drain with large amounts of water. Be sure that HF does not come in contact with the skin during this process–use gloves!

3. We do not recommend the use of perchloric acid ($HClO_4$) in the school laboratory because of the difficulty of handling it safely. If you have some in storage, keep it separated from other chemicals, labeled correctly, and record that it is to be discarded by appropriate means. Do not dispose of it yourself.
4. Bases can be neutralized by forming the aqueous solution, followed by neutralization with sodium hydrogen sulfate ($NaHSO_4$, also known as sodium bisulfate). Once neutralized, they may be washed down the sink with large amounts of water.
5. Procedures for dealing with the disposal of salts of heavy metals are addressed on p. 21.
6. Organic solvents, such as alcohol, hexane, *etc*. can be effectively disposed of by allowing the solvent to slowly evaporate from an open large-mouth container or flat dish in a chemical hood. It is most effective if a large beaker or jar is used and if a paper towel is used as a wick in the container to increase the surface from which the material can evaporate. Since the school is a very small generator of waste, this is an acceptable practice for that type of laboratory.

Note: We urge the instructor to carefully consider alternatives to the use of experiments in which organic solvents play a major role. In particular, the use of chlorinated solvents is discouraged, due to their adverse effect on the environment, and the teacher should use suitable alternatives, if at all possible.

7. Oxidants, such as permanganate or chromate ion have two possible hazards: they contain heavy metal toxins and are potential carcinogens, and they are strong oxidants that must be kept at a discreet distance from potential reductants. They may be reduced to lower oxidation states by the addition of sodium hydrogen sulfite. Such a treatment is required before the metals are precipitated as indicated above.
8. Flammable materials, such as organic solvents, must both be stored and disposed of in a manner such that they are kept from potential sources of ignition and/or chemical oxidants.
9. Active metals must be sealed in air-tight containers to prevent their contact with water or strong oxidants, with which they will react to produce sufficient heat to ignite combustible materials.

Who Owns It?

Once a supplier has sold and shipped a chemical to the teacher, and once it has been received, it belongs to the school until one of the following happens:

1. It is consumed in an experiment.

2a. It is converted to a product in a reaction, in which case the product becomes the waste of concern;

2b. It is deactivated by being converted to an innocuous product that can be disposed of by means of steps 1, 3 or 4.

3. It is disposed of by proper means in the laboratory.
4. It is disposed of by transportation to an acceptable and approved chemical waste disposal facility.

The chemical is essentially the property of the school until it is in its final resting place or in its final disposition. In the case of disposal to a chemical dump, of course, regulations of the EPA and Department of Transportation pertain with regard to packaging, labeling, and transportation. The teacher's responsibility is to provide complete documentation about the contents of packages, to pack the shipment containers correctly, and to provide the information required for bills of lading, *etc.*

When you get to this point in deciding to dispose of materials you will need specific advice from a knowledgeable person, and you will benefit from the information in the pamphlet *Safety in the Academic Laboratory*, 5th edition, *(1)*.

Produced by the High School Subcommittee, Joint Board/Council Committee on Chemical Safety, American Chemical Society: Tomás G. Berger B., K. O. Berry,* W. H. Breazeale, Jr., D. Mandt, E. K. Mellon.

*Deceased, March 30, 1993.

APPENDIX I

EXTREMELY HAZARDOUS CHEMICALS

(Note: ***Normally none of these chemicals should be present in the high school laboratory. The safety hazard involved generally outweighs any pedagogical value they may possess.*** Refer to definitions of terms found at the end of this Appendix.)

SUBSTANCE	CAS NO.	HAZARD	COMMENT
2-Acetylaminofluoride	53-96-3	Regulated human carcinogen	Poison B
4-Aminobiphenyl	92-67-1	Regulated human carcinogen	Poison B
Aniline	142-04-1	Acute Toxin	Poison B
Arsenic, powder	7440-28-2	Poison	Poison B
Arsenic trioxide	1327-53-3	Acute toxin	Poison B
Asbestos (generic term for substances such as: antophyllite; chrysolite; tremolite)	1332-21-4 17068-78-9 12001-29-5 14567-73-8	Regulated human carcinogen	ORM-E
Barium peroxide	1304-29-6	Powerful oxidizer	Oxidizer
Benzene	71-43-2	Flammable, acute toxin, suspected carcinogen	Poison B
Benzidine (and salts)	92-87-5	Regulated human carcinogen	Flammable
Benzoyl peroxide	94-36-0	May explode by heat, shock, or friction	Oxidizer
Cadmium, powder	7440-43-6	Poison. Dust is known carcinogen	ORM-E
Carbon disulfide	75-15-0	Acute toxin, small amounts in air may be explosive	Flammable
Carbon tetrachloride	56-23-5	Narcotic, suspected carcinogen	ORM-A

SUBSTANCE	CAS NO.	HAZARD	COMMENT
Chloroform	67-66-3	Narcotic, suspected carcinogen	ORM-A
Chromium (VI) oxide	333-82-0	Suspected carcinogen	Oxidizer
3,3-Dichlorobenzidine	91-94-1	Regulated human carcinogen	Poison B
Diisopropyl ether		If more than 12 months old, one should assume peroxides have formed	Flammable
Dimethylamine	62-75-9	Acute toxin	Flammable
4-Dimethylaminoazobenzene	60-11-7	Regulated human carcinogen	Poison B
Ethylene dichloride	569-57-3	Suspected human carcinogen	Flammable
Ethyleneimine	151-56-4	Regulated human carcinogen, flammable, poison	Flammable
Ethylene oxide	75-21-8	Suspected human carcinogen, acute toxin	Flammable
Hydrazine (anhydrous)	302-01-2	Suspected human carcinogen, flammable, acute toxin	Flammable
Hydrofluoric acid	7664-39-3	Corrosive, toxic	Corrosive
Lead arsenate	7784-49-9	Highly toxic	Poison B
Methylchloromethyl ether	107-30-2	Regulated human carcinogen	Flammable
4,4-Methylene-bis-(2-chloroaniline)		Regulated human carcinogen	Poison B
Methylene chloride	75-09-2	Suspected human carcinogen, acute toxin	ORM-A
α-Naphthylamine	134-32-7	Regulated human carcinogen	Poison B

SUBSTANCE	CAS NO.	HAZARD	COMMENT
β-Naphthylamine	91-59-8	Regulated human carcinogen	Poison B
Nickel, powder	7440-02-0	Acute toxin	Metal
Nicotine	541-11-5	Acute toxin	Poison B
4-Nitrobiphenyl	92-93-3	Regulated human carcinogen	Poison B
Nitrogen triiodide		Explosive when dry	Explosive A
Perchloric acid		Strongly reacts with organics	Oxidizer
Phosphorus (yellow/white)	7723-14-0	Deteriorates with age	Flammable solid
Picric acid	88-89-1	Dry picric acid is explosive and can be detonated by shock or heat	Explosive A
Potassium metal	7440-09-7	Becomes more dangerous with age, forms peroxides	Metal
β-Propiolactone	57-57-8	Regulated human carcinogen	Poison B
Sodium arsenate	7631-89-2	Suspected carcinogen, acute toxin	Poison B
Sodium arsenite	7784-46-5	Suspected carcinogen, acute toxin	Poison B
Sodium azide	26628-22-8	Unstable and explosive	Explosive B
Vinyl chloride	75-01-4	Regulated human carcinogen	Flammable solid

DEFINITIONS OF TERMS USED IN THE 'COMMENT' COLUMN OF THIS APPENDIX

Explosive, Class A	Liquid or gaseous explosive which functions by detonation.
Explosive, Class B	Explosives which function by rapid combustion.

Explosive, Class C	Contains restricted quantities of Class A and Class B explosives.
Compressed gas	Material having an absolute pressure exceeding 40 pounds per square inch at 70 °F or an absolute pressure exceeding 104 psi at 130 °F, or any liquid flammable material having a vapor pressure exceeding 40 psi at 100 °F.
Corrosive	A liquid or solid that causes visible destruction or reversible alterations in human skin at the site of contact or a liquid that has a severe corrosion rate on steel.
Combustible Liquid	Has flash point between 37.8 °C and 93.3 °C.
Flammable gas	Any material that meets the definition of compressed gas.
Flammable liquid	Any liquid having a flash point below 37.8 °C.
Flammable solid	Any solid material other than an explosive which is liable to cause fires through friction or which can be ignited readily.
Metal	Active metal, strong reductant.
Oxidizer	Reacts with strong reductant, producing heat and potentially explosive products.
Poison, Class A	Poisonous gases or liquids of such a nature that a very small amount of the gas, or vapor of the liquid, mixed with air, is dangerous to life.
Poison, Class B	Those substances, liquid or solid, other than Class A poisons or irritating materials, which are known to be so toxic to humans as to afford a hazard to health during transportation.
Acute	Extreme sensitivity, producing its effect with small amounts or one-time application.

ORM - Other Regulated Materials

ORM, Class A	Irritating or noxious substances.
ORM, Class B	Corrosive that dissolves 1/4" aluminum per year.
ORM, Class C	Any substance, not otherwise regulated, that has inherent properties that require special handling.
ORM, Class D	Substance with limited hazards.
ORM, Class E	Substance not included in any hazard class that is commonly in the form of hazardous waste.

APPENDIX II

ABBREVIATED LIST OF CHEMICALS THAT MAY BE PRESENT IN HIGH SCHOOLS

LEGEND FOR CODES

CAS NO - Chemical Abstract Number
NFPA - National Fire Protection Association

CODES

[A] Those that are known carcinogens, mutagens, or teratogens. We generally do not recommend using these chemicals.

[B] Those that are suspected carcinogens, mutagens, or teratogens. We recommend that use of these chemicals be greatly restricted.

[C] Found in Flinn's list of "40 Devils" (5), which represent routine sources of trouble in the high school laboratory. Use of these chemicals should also be avoided as much as possible.

[D] Those that pose sufficient risk, other than those listed above, that they probably should not be in the high school laboratory. We generally do not recommend using these chemicals.

[E] Those that may be used by a qualified chemistry teacher, but probably should not be used by students.

[F] Those that have potential risk associated with them, but whose use by students may be justified, providing the smallest amounts possible are used and the students have been informed as to the correct means of handling these chemicals.

(**Note:** Chemicals that were included in Appendix I do not also appear in this list.)

Compound Name	CAS Number	Codes	NFPA Ratings Health Flammability Reactivity	Comments
Acetamide	60-35-5	B		
Acetic acid	64-19-7	C		Corrosive; flammable liquid
Acetone	67-64-1	F		Extremely flammable
Acetyl chloride	73-36-5	C		Lachrymator; violent reaction with water
Acetylene	74-86-2	D	1 4 3	Asphyxiant
Acid Green	4680-78-8	D		Very toxic; biological stain
Acrylonitrile	107-13-1	A D	4 3 2	Dust especially bad
Adipoyl chloride	111-50-2	C E		Corrosive
Aluminum, metal	7429-90-5	F		Powder or dust may explode
Aluminum chloride, anhydrous	7446-70-0	C E	3 0 2	Respiratory irritant; violent reaction with water; corrosive
Ammonia, aqueous, concentrated	7664-41-7	C F		Respiratory, skin and eye hazard; corrosive
Ammonium chromate	7788-98-9	A D		Dust is carcinogen; difficult to prepare for disposal
Ammonium dichromate	7789-09-5	B E	1 1 1	Fire hazard; dust is suspected carcinogen
Ammonium metavanadate	7803-55-6	F		Very toxic
Ammonium nitrate	6484-52-2	F		Decomposes above 160 °C, producing noxious and explosive gases
Ammonium oxalate	6009-70-7	F		Toxic by inhalation

Compound Name	CAS Number	Codes	NFPA Ratings Health Flammability Reactivity	Comments
Ammonium persulfate	7727-54-0	B F		Oxidizer
Aniline hydrochloride	142-04-1	B E		Acute toxin
Anthracene	120-12-7	E	0 1 0	Toxic; respiratory irritant
Antimony	7440-36-0	E		Fumes are toxic; compounds can be absorbed through skin
Antimony(III) oxide	1390-64-4	E		Very toxic
Antimony pentachloride		C D		Corrosive; reacts with organics
Antimony potassium tartrate (potassium antimonyl tartrate hemihydrate)	28300-74-5	F		Toxic by inhalation
Antimony trichloride	10025-91-9	C D		Corrosive; toxic
Arsenic(V) oxide	1303-28-2	A D		Highly toxic
Arsenic trichloride	7784-34-1	B D	3 0 0	Highly toxic
Ascarite	81133-20-2	A		Caustic; NaOH coated asbestos
Auramine	2465-27-2	F		Biological stain; toxic
Barium chloride	10361-37-2	F		Very toxic as heavy metal by ingestion. (All soluble compounds containing Ba are very toxic.)
Barium oxalate	516-02-9	E		Very toxic
Basic fuchsin	569-61-9	A F		
Benzyl alcohol	100-51-6	C D		Strong reductant

Compound Name	**CAS Number**	**Codes**	**NFPA Ratings Health Flammability Reactivity**	**Comments**
Beryllium carbonate	13106-47-3	B D		Highly toxic
Bis-Chloromethylether	542-88-1	A		Potent carcinogen; very toxic
Bromine	7726-95-6	C E		Highly toxic by ingestion; strong oxidant, severe skin irritant; reacts violently with organics
Butyl alcohols *n*-butyl *t*-butyl	71-36-3 75-65-0	E		Flammable; prolonged inhalation can be toxic; eye irritant; absorbed by skin
Butyric acid	107-92-6	C D		Stench; lachrymator
Cadmium acetate	89759-80-8	A D		When in a dilute aqueous solution, hazard is likely only on ingestion - a condition true for all cadmium salts
Cadmium bromide	7789-42-6	B E		
Cadmium carbonate	513-78-0	B E		
Cadmium chloride	10108-64-2	A F		
Cadmium sulfate	10124-36-4	A F		
Calcium carbide	75-20-7	C E	1 4 2	Flammable; reacts with water to give acetylene; fire risk
Calcium fluoride	7789-75-5	F		Toxic; skin irritant; reacts with water to produce hydrofluoric acid
Calcium hydroxide	1305-62-0	F		Caustic; corrosive base
Calcium hypochlorite	7778-54-3	D E		Corrosive powder; harmful to eyes, lungs, and skin
Calcium oxide (lime)	1305-78-8	F		Corrosive base

Compound Name	CAS Number	Codes	NFPA Ratings Health Flammability Reactivity	Comments
Carmine	1260-17-9	F		Biological stain
Catechol	120-80-9	E		Toxic, skin irritant; respiratory irritant
Cellulose nitrate	9004-70-0	E	2 3 3	Explosive
Chloral hydrate	302-17-0	D		Hypnotic drug
Chloretone (1,1,1-trichloro -2-methyl-2-propanol)	6001-64-5	D		Hypnotic drug
Chlorine	7782-50-5	E	3 0 0	Pungent, toxic gas
Chlorobenzidines		D		
Chloroprene	126-99-8	D	2 3 0	Very toxic
Chromates		A E		Eye, skin, respiratory irritant
Chromium	7440-47-3	A F		Dust is a carcinogen; compounds are toxic by ingestion
Chromium(II) acetate	4112-22-5	E		Easily oxidized
Chromium(III) acetate	1066-30-4	F		Toxic
Chromium(III) oxide	138-38-9	F		Toxic
Chromium(III) potassium sulfate	7788-99-0	F		Toxic by ingestion
Cobalt	7440-48-4	F		Dust is toxic
Cobalt nitrate hexahydrate	10026-22-9	F	1 0 0	Dry solid is oxidizer; toxic
Colchicine	64-86-8	D		Highly toxic

Compound Name	CAS Number	Codes	NFPA Ratings Health	Flammability	Reactivity	Comments
Copper salts		F				Toxic by ingestion
Cyclohexane	110-82-7	F	1	3	0	Flammable; toxic by inhalation and contact
Cyclohexene	110-83-8	F	1	3	0	Flammable; toxic by inhalation
p-Dichlorobenzene	106-46-7	B E	2	2	0	Severe irritant; toxic by ingestion
1,2-Dichloroethane (Ethylene dichloride)	07-06-2	A D	2	3	0	
Dichloromethane (Methylene chloride)	75-09-2	F	2	1	0	Narcotic in high concentration
Dichlorophenol	620-45-1	F				
Diethyl sulfate (Ethyl sulfate)	64-67-5	A D	3	1	1	
Dinitrogen tetroxide	10102-44-0	E	3	0	0	Extremely toxic
N,N-Dimethylaniline	21-69-7	E	3	2	0	Toxic
Dimethyl sulfate	77-78-1	A D	4	2	0	
2,4-Dinitrophenol	51-28-5	F				Very toxic; flammable
p-Dioxane	123-91-1	A C D	2	3	1	Forms peroxides on standing for long periods
Diphenylamine	122-39-4	A E	3	1	0	
Diphenyl carbonate	102-09-0	E				Toxic
Ethyl ether (Diethyl ether, Ether, Diethyl oxide)	60-29-7	C E	2	4	1	Extremely flammable; forms peroxides on standing

Compound Name	CAS Number	Codes	NFPA Ratings			Comments
			Health	Flammability	Reactivity	
Formaldehyde (Formalin)	50-00-0	B C E	2	2	0	Strong irritant
Formic acid	64-18-7	E				Toxic; corrosive; conc. form subject to explosion
Fuchsin acid	244-88-0	F				Plasma stain
Gasoline	8006-61-9	A E	1	3	0	Contains benzene; fumes are toxic; very flammable; deteriorates on storage
Gunpowder		E				Flammable solid mixture
Hematoxylin	517-28-2	B E				Biological stain
Hexachlorophene	70-30-4	A D				
Hydroiodic acid (Hydriodic acid)	10034-85-2	F	3	0	0	Corrosive; light sensitive
Hydrobromic acid	10035-10-6	F	3	0	0	Corrosive
Hydrochloric acid	7647-01-0	F	3	0	0	Corrosive
Hydrogen gas	1333-74-0	E				Explosive
Hydrogen peroxide, 30%	7722-84-1	C E				Strong catalyst; causes burns; strong oxidant
Hydrogen sulfide	7783-06-4	D	3	4	0	Extremely toxic
Hydroquinone	123-31-9	F	2	1	0	Corrosive; toxic by ingestion and inhalation
Indigo carmine	860-22-0	F				Light sensitive dye
Iodine	7553-56-2	F				Vapor corrosive to eyes and respiratory tract
Iodomethane	74-88-4	A D				

Compound Name	CAS Number	Codes	NFPA Ratings Health Flammability Reactivity	Comments
Isoamyl alcohol (3-Methyl-1-butanol)	123-51-3	F	1 2 0	Vapor is irritating
Isobutyl alcohol (2-Methyl-1-propanol)	78-83-1	F	1 3 0	Vapor is irritating
Kerosene		F		Flammable; skin irritant
Isopropyl alcohol	26232-83-7	F	1 3 0	Flammable
Lead	7439-92-1	B E		Dust is very harmful
Lead(II) acetate	301-04-2	A F		
Lead(II) carbonate, basic	598-63-0	E		Very toxic
Lead(II) chloride	7758-97-6	F		Highly toxic
Lead(II) nitrate	10099-74-8	F	1 0 0	Oxidant; toxic by ingestion or inhalation
Lead(IV) oxide	1309-60-0	F		Strong oxidizer; toxic
Lithium, metal	7439-93-2	C E	1 1 2	Toxic metal; flammable; caustic and corrosive
Lithium chloride	7447-41-8	F		Hygroscopic
Lithium nitrate	7790-69-4	F		Toxic
Magenta I (Basic Fuchsine)	569-61-9	A E		
Magnesium, metal (powder; ribbon)	7439-95-4	C E F	0 1 2	Flammable solid; moisture sensitive combustible solid
Magnesium perchlorate	10034-81-8	C D	1 0 0	Oxidizer; explosive; irritant; unstable
Manganese dioxide	1313-13-9	F		Strong oxidizer; toxic

Compound Name	**CAS Number**	**Codes**	**NFPA Ratings Health Flammability Reactivity**	**Comments**
Mercury, metal	7439-97-6	C E		Highly but subtly toxic
Mercury compounds				All mercury compounds are toxic; harmful by swallowing, inhaling, or skin absorption
Mercury(I) chloride (calomel)	10112-91-1	E		Toxic
Mercury(II) chloride	7487-94-7	E		Highly toxic
Mercury(I) nitrate	7782-86-7	F		Highly toxic
Mercury(II) nitrate	7783-34-8	F		Highly toxic; short shelf life
Mercury(I) oxide	15829-3-5	E		Highly toxic
Mercury(II) oxide	21908-53-2	E		Highly toxic; light sensitive
Mercury(II) sulfate	7783-353-9	E		Highly toxic; light sensitive
Mercury(II) sulfide	1344-48-5	E		Toxic; light sensitive
Mesitylene (1,3,5-Trimethylbenzene)	108-67-8	E	0 2 0	Flammable
Methane, gas		F		Flammable gas; toxic
Methanol	67-56-1	F		Flammable; very toxic
Methyl ethyl ketone (2-Butanone)	78-93-3	F	1 3 0	Flammable liquid; narcotic by inhalation
Methyl iodide (Iodomethane)	74-88-4	A D		
Methyl methacrylate	0-62-2	D	2 3 2	Explosive; powder is respiratory irritant
Methyl oleate	112-62-9	F		Irritant; some salts are poisonous

Compound Name	CAS Number	Codes	NFPA Ratings Health	Flammability	Reactivity	Comments
Methyl Orange	547-58-0	F				
Methyl Red	493-52-7	F				
Nickel(II) acetate	6018-89-9	A F				Dust is carcinogen; however, dilute aqueous solutions are believed safe for student use
Nickel ammonium sulfate	15699-18-0	A F				Dust is carcinogen; see preceding note
Nickel(II) carbonate	3333-67-3	A F				Dust is carcinogen; see preceding note
Nickel(II) chloride	7791-20-0	A F				Dust is carcinogen; see preceding note
Nickel(II) nitrate	13478-00-7	A F	1	0	0	Dust is carcinogen; see preceding note
Nickel(II) oxide	1313-99-1	A F				Dust is carcinogen; see preceding note
Nitric acid	7697-37-2	C F				Strong oxidizer; noxious fumes; decomposes over time
Nitrocellulose	9004-70-0	E	2	3	3	Explosive
Nitrogen dioxide (Dinitrogen tetroxide)	10102-44-0	E	3	0	0	Very toxic
2-Nitronaphthalene	581-89-5	B D	1	1	0	
Nitrosamines		A D				
N-Nitrosodimethylamine	62-75-9	A D				
4-Nitrosophenol	104-91-6	A D				
Osmium tetroxide (osmic acid)	20816-12-0	E				Extremely toxic
Oxalic acid	6153-56-6	F				Toxic when swallowed or absorbed through skin

Compound Name	**CAS Number**	**Codes**	**NFPA Ratings Health Flammability Reactivity**	**Comments**
Oxygen gas	7782-44-7	F		Compressed gas; supports combustion
Paraffin wax	8002-74-2	F		Fumes are flammable; toxic
Paris Green	12002-03-8	D		Highly toxic; contains arsenic
Pentane	109-66-0	F	1 4 0	Highly flammable
Petroleum ether (Ligroin)	8032-32-4	F	1 4 0	Highly flammable
Phenol	108-95-2	B E	3 2 0	Highly toxic; corrosive; absorbed through skin
Phenylthiourea (phenylthiocarbamide)	103-85-5	D		Highly toxic; decomposes to form sulfur and nitrogen oxide fumes
Phosphorus, red	7723-14-0	F	0 1 1	Flammable solid; less reactive than white phosphorus
Phthalic anhydride	85-44-9	F	2 1 0	Toxic
Polybrominated biphenyls		A D		
Polychlorinated biphenyls		A D		
Polychlorodibenzofurans		A D		
Potassium chlorate	3811-04-9	E	2 0 0	Irritant; strong oxidizer; may explode when shocked or heated
Potassium chromate	7789-00-6	A F		Toxic
Potassium dichromate	7778-50-9	B F		Toxic
Potassium ferricyanide	13746-66-2	F		
Potassium ferrocyanide	14459-95-1	F		

Compound Name	CAS Number	Codes	NFPA Ratings Health Flammability Reactivity	Comments
Potassium hydroxide	1310-58-3	F		Toxic; caustic, corrosive
Potassium oxalate	6487-48-5	F		Hygroscopic; toxic
Potassium periodate	7790-21-8	F		Irritant; strong oxidizer
Potassium permanganate	7722-64-7	C F	1 0 0	Skin irritant; powerful oxidizer; explodes on sudden heating
Potassium sulfide	1312-73-8	F	2 1 0	Fire risk; produces H_2S with acids
Potassium thiocyanate	333-20-0	F		
Propane		D E		Flammable gas; narcotic in high concentrations
1-Propanol (*n*-propyl alcohol)	71-23-8	F	1 3 0	Flammable liquid
2-Propanol (isopropyl alcohol)	26232-83-7	F	1 3 0	Flammable liquid
Pyridine	110-86-1	E	2 3 0	Flammable liquid; irritant
Pyrogallic acid (pyrogallol)	87-66-1	E		Toxic irritant
Saccharin (*o*-Benzoic sulfimide)	81-07-2	B		
Salicylamide (2-Hydroxybenzamide)	65-45-2	E		Emits highly toxic fumes when heated
Salol (Phenyl salicylate)	118-55-8	F		Toxic by ingestion
Sebacoyl chloride	111-19-3	C E		Respiratory and eye irritant; corrosive
Selenium	7782-49-2	E		Toxic

Compound Name	CAS Number	Codes	NFPA Ratings Health Flammability Reactivity	Comments
Silver nitrate	7761-88-8	F	1 0 0	Corrosive solid; toxic; undergoes photodecomposition
Sodium, metal	7440-23-5	C F	3 1 2	Flammable solid; reacts violently with water producing hydrogen
Sodium bromate	7789-38-0	E		Powerful oxidizing agent; toxic; skin irritant
Sodium chlorate	7775-09-9	E	1 0 2	Oxidizer; corrosive
Sodium chromate	1034-82-9	B F		Dust is suspected carcinogen; toxic
Sodium dichromate	10588-01-9	B E	1 0 1	Highly toxic; oxidizer; dust is suspected carcinogen
Sodium ferrocyanide	13601-19-9	F		
Sodium fluoride	7681-49-4	F	2 0 0	Toxic by ingestion and inhalation; strong skin irritant
Sodium hydroxide	1310-73-2	F		Strong irritant; caustic; poisonous, corrosive
Sodium hypochlorite	7681-52-9	F		Caustic, poisonous, irritating to skin and respiratory tract
Sodium nitrate	7631-99-4	F	1 0 0	Strong oxidizer when dry or molten
Sodium nitrite	7632-00-0	F		Oxidizer; produces nitrogen dioxide
Sodium oxalate	62-76-0	F		Toxic by ingestion; short shelf life
Sodium peroxide		C D		Explosion/fire risk in combination with powdered metals and organics; oxidizer
Sodium silicate		F		Skin and mucous membrane irritant
Sodium silicofluoride	16893-85-9	E		Leads to silicosis, causing death or permanent injury after short exposure to small quantities

Compound Name	CAS Number	Codes	NFPA Ratings			Comments
			Health	Flammability	Reactivity	
Sodium sulfide	1313-82-2	F	2	1	0	Flammable solid; corrosive
Sodium thiosulfate	540-72-7	F				Toxic irritant
Stannic chloride (Tin(IV)chloride)	7646-78-8	F	3	0	1	Toxic; corrosive
Stannous chloride (Tin(II)chloride)	7772-99-8	F				Toxic
Stearic acid	54-11-4	F	1	1	0	
Strontium, metal	7440-24-6	E				Flammable solid when powdered
Strontium nitrate	10042-76-9	F	1	0	0	
Styrene	100-42-5	E				Flammable
Sudan III	85-86-9	F				Biological stain
Sudan IV	85-83-6	B F				
Sulfamethazine	57-68-1	D				Decomposes when heated, producing sulfur oxides
Sulfuric acid	7664-93-9	C F	3	0	2	Strong acid; very corrosive; skin irritant
Sulfuric acid, fuming	7664-93-9	D	3	0	2	Strong acid; very corrosive; skin irritant
Talc (hydrous magnesium silicate)	14807-96-6	F				May contain asbestos
Tannic acid	1401-55-4	B E	0	1	0	
1,1,2,2-Tetrabromoethane (Acetylene tetrabromide)	79-27-6	B D	3	0	1	
Thallium compounds		E				Extremely toxic
Thermit	8049-32-9	E				Fire hazard

Compound Name	CAS Number	Codes	NFPA Ratings Health Flammability Reactivity	Comments
Thioacetamide	62-55-5	A F		Suspected carcinogen
Thiourea	62-56-6	B E		
Titanium trichloride	7705-07-9	E		Fire risk; hydrolyzes to produce HCl vapor
Toluene	108-88-3	F	2 3 0	Flammable liquid; toxic by ingestion, inhalation or skin absorption
o-Toluidine	95-53-4	B E	3 2 0	
Trichloroethylene	79-01-6	B E	2 1 0	
1,1,2-Trichloro-1,2,2-trifluoroethane (TTE)	79-01-6	F		Skin irritant
Turpentine		F		Flammable; skin irritant
Uranium, metal	7440-61-1	D		Radioactive; heavy metal
Uranyl acetate	541-09-3	F		Highly toxic
Uranyl nitrate	10102-06-4	F	1 0 0	Oxidizer; highly toxic
Urethane (Ethyl carbamate)	51-79-6	B E		
Vinyl chloride	75-01-4	A D		Flammable
Wood's metal	76093-98-6	F		Toxic; lead alloy
Xylenes, mixed	1330-20-7	F	2 3 0	Flammable liquid
Zinc, metal	7440-66-6	F		Powder is combustible

Appendix III

Comparison of J.T. Baker and NFPA Systems to Express Safety of Chemicals

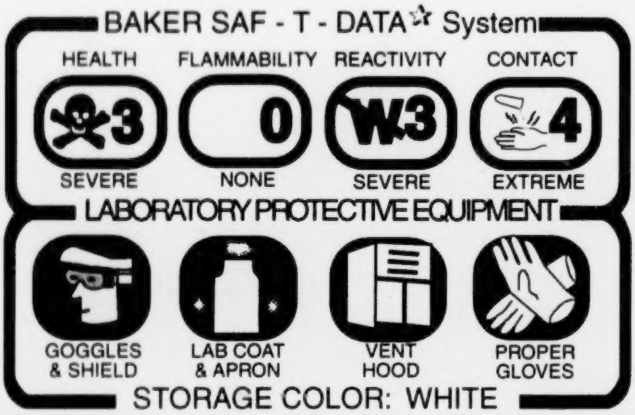
The chemical labels required by the National Fire Protection Association (NFPA) in their Code 704, to which most schools are expected to conform, were developed as a means by which firefighters could be informed about the content of a room, laboratory, or bottle and the inherent dangers in it. They were not meant to be codes by which chemists communicate information. For that reason, they are often inadequate to communicate chemical information to employees and students.

A comparison of code systems reveals that the NFPA system is not only inadequate, but it is often very difficult to use when trying to define codes for the hundreds or thousands of chemicals for which codes do not already exist. A comparison of NFPA codes and those of the J. T. Baker Saf-T-Data system is included on the following pages. Other chemical companies and suppliers of chemicals may have developed alternate coding systems. The NFPA health ratings and codes are qualitative rather than quantitative and are divided into five categories as compared to eight categories in the Baker system.

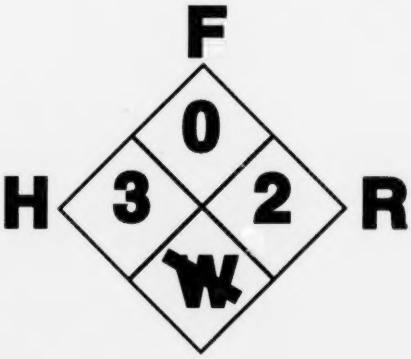
While the flammability codes and parameters are approximately the same for the two code systems, problems arise in comparisons of "reactivity." The NFPA system defines "reactivity" as thermodynamic stability. This is based on a firefighter's concern whether a chemical is stable or whether it may explode. The NFPA carefully stipulates that their definition of reactivity implies that substances are reactive "in themselves." This does not conform with the general term "chemical reactivity." For instance, there are very few compounds found in the school laboratory which would have a *stability* rating above 1 (see the definitions), whereas there are likely to be strong oxidizers, for instance, that would have a *chemical reactivity* rating of 4. For chemists and chemistry students, we conclude that chemical reactivity is much more meaningful and communicates more useful information to the students.

Whereas the NFPA system indicates a contact hazard, it does not quantify the degree of the risk. The Baker system includes a rating system. We believe the Baker system, or a similar system such as the Flinn "At-a-Glance™" system, includes more useful information for the school laboratory than the NFPA system, although it may be best to use both systems in parallel.

BAKER SAF-T-DATA SYSTEM


BAKER SAF - T - DATA☆ System
HEALTH
FLAMMABILITY
REACTIVITY
CONTACT
3
0
W3
4
SEVERE
NONE
SEVERE
EXTREME
LABORATORY PROTECTIVE EQUIPMENT
GOGGLES & SHIELD
LAB COAT & APRON
VENT HOOD
PROPER GLOVES
STORAGE COLOR: WHITE

NFPA SYSTEM



LABEL FOR SULFURIC ACID

HAZARD RATINGS

- **Health Hazard** - the danger or toxic effect a substance presents if inhaled, ingested or absorbed.
- **Flammability Hazard** - the tendency of the substance to burn.
- **Reactivity Hazard** - the potential of a substance to explode or react violently with air, water or other substances.
- **Contact Hazard** - the danger a substance presents when exposed to skin, eyes and mucous membranes.

RATING SCALE

0 None (or no scientific data in the standard references that suggest the substance is hazardous)

Slight	Moderate	Severe	Extreme
1	2	3	4

STORAGE COLOR CODES

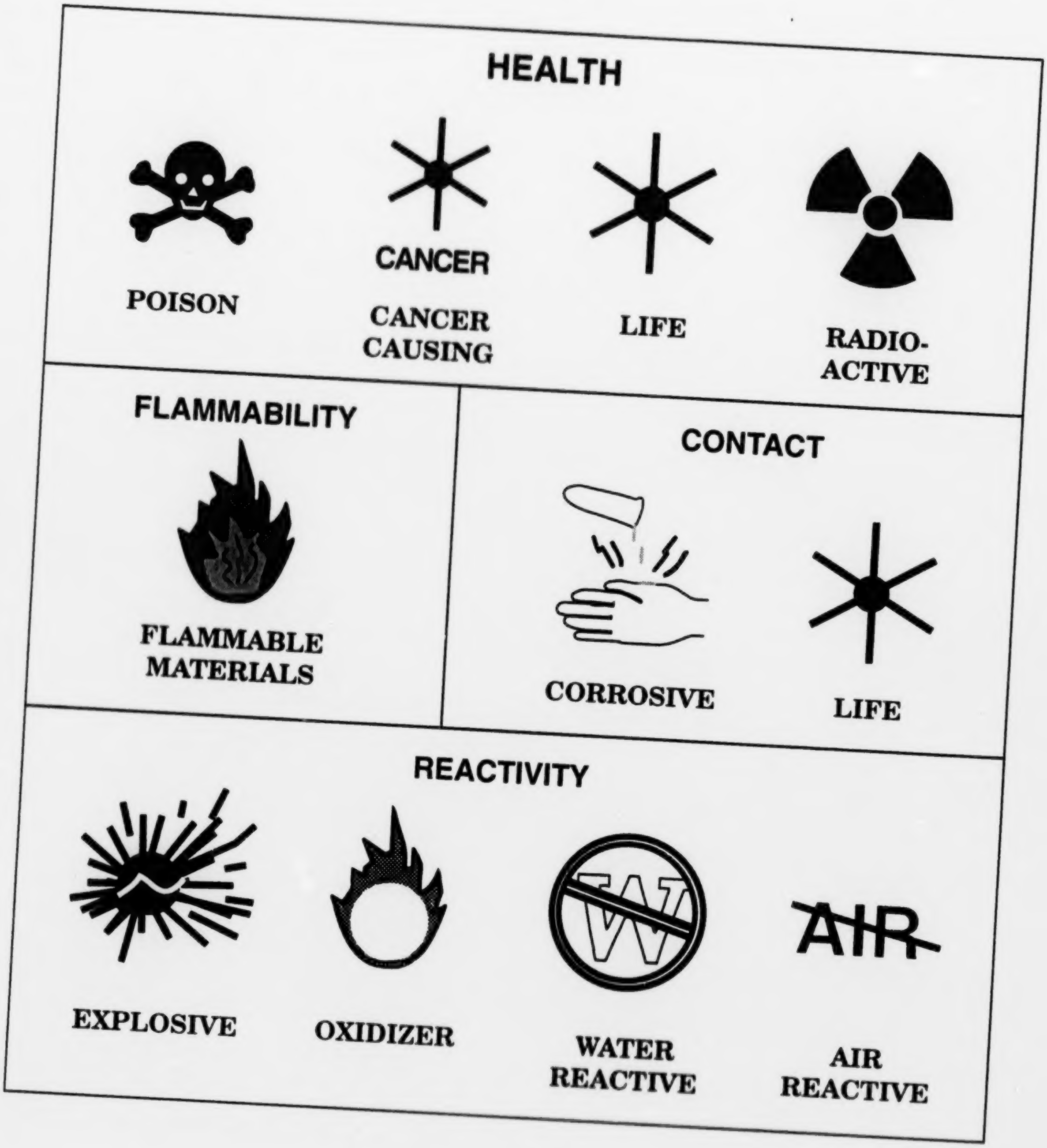
- BLUE - Severe health risk
- RED - Flammable
- YELLOW - Highly reactive
- WHITE - Corrosive
- ORANGE - Minimal storage risk
- STRIPED - Special hazard

To protect the professional fire fighter, NFPA (National Fire Protection Association) has established a numerical code that rates chemicals *under fire conditions* in four categories:

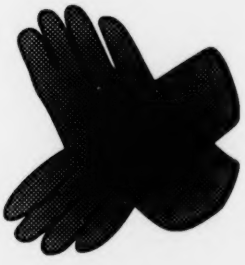
H - Health
F - Flammability
R - Reactivity

and **unusual reactivity**. Within each category, a numerical rating system has been established in five numerical ratings (0-4). Number 4 is a severe hazard, and number 0 is no special hazard. A few state laws require this rating system on chemical substance labels. The NFPA numerical ratings exist for a very limited number of substances and the numbers represent hazard *under fire conditions* as opposed to normal laboratory use, a situation that tends to distort hazard characteristics. Teachers are urged to depend on other systems and on the safety pictorials shown elsewhere on each label.

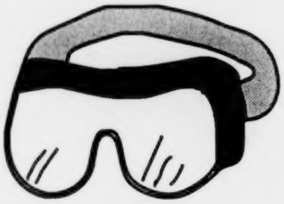
HAZARD PICTOGRAMS


HEALTH
CANCER
POISON
CANCER CAUSING
LIFE
RADIO-ACTIVE
FLAMMABILITY
FLAMMABLE MATERIALS
CONTACT
CORROSIVE
LIFE
REACTIVITY
W
AIR
EXPLOSIVE
OXIDIZER
WATER REACTIVE
AIR REACTIVE

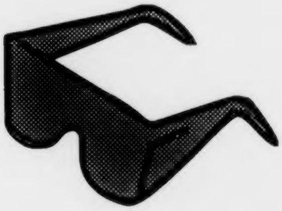
LABORATORY PROTECTIVE DEVICES



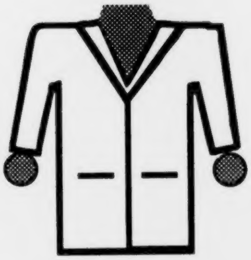
PROPER GLOVES



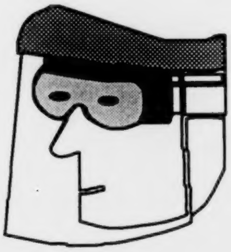
GOGGLES



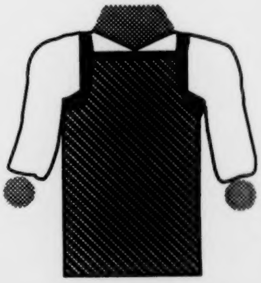
SAFETY GLASSES



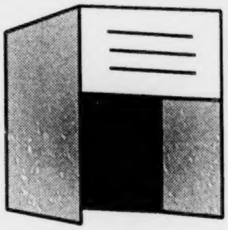
LAB COAT



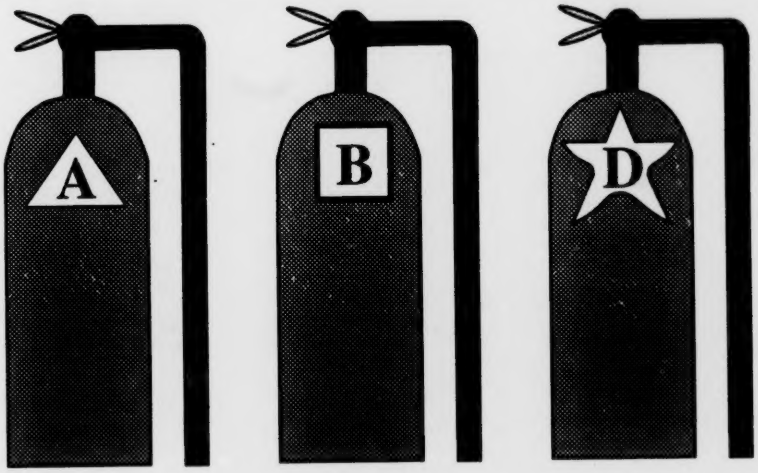
GOGGLES & SHIELD



LAB COAT & APRON



VENT HOOD



EXTINGUISHERS

Appendix IV

FOOTNOTES AND BIBLIOGRAPHIC REFERENCES

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Appendix VII

Minimum Safety Guidelines for Chemical Demonstrations

ACS Division of Chemical Education

Chemical Demonstrators *Must*:

1. know the properties of the chemicals and the chemical reactions involved in all demonstrations presented.
2. comply with all local rules and regulations.
3. wear appropriate eye protection for all chemical demonstrations.
4. warn the members of the audience to cover their ears whenever a loud noise is anticipated.
5. plan the demonstration so that harmful quantities of noxious gases (e.g., NO_2, SO_2, H_2S) do not enter the local air supply.
6. provide safety shield protection wherever there is the slightest possibility that a container, its fragments, or its contents could be propelled with sufficient force to cause personal injury.
7. arrange to have a fire extinguisher at hand whenever the slightest possibility for fire exists.
8. **not** taste or encourage spectators to taste any non-food substance.
9. **not** use demonstrations in which parts of the human body are placed in danger (such as placing dry ice in the mouth or dipping hands into liquid nitrogen).
10. **not** use "open" containers of volatile, toxic substances (e.g., benzene, CCl_4, CS_2, formaldehyde) without adequate ventilation as provided by fume hoods.
11. provide written procedure, hazard and disposal information for each demonstration whenever the audience is encouraged to repeat the demonstration.
12. arrange for appropriate waste containers for and subsequent disposal of materials harmful to the environment.

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

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